

Calculations In Chemistry



Module 15 – Redox Reactions

Module 16 – Half-Reaction Balancing



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Module 15 — Redox Reactions

Timing: Many courses cover *redox reactions* as part of a unit on molarity and solution reactions. Others cover redox reactions as part of an electrochemistry unit later in the course. Complete Module 15 at the time that redox reactions are assigned in *your* class.

Balancing redox reactions using *oxidation numbers* is covered in this module. Balancing redox using *half-reactions* is covered in Module 16. If both balancing using oxidation numbers and half-reactions are assigned now, complete both modules. If half-reactions are assigned at the later time than balancing with oxidation numbers, complete Module 16 when half-reaction balancing is assigned.

* * * * *

Lesson 15A: Oxidation Numbers

Pretest: If you can solve these problems, you may skip to Lesson 15B. Answers are at the end of the lesson.

1. Define oxidation.
2. What is the oxidation number of
 - a. Each chromium atom in $K_2Cr_2O_7$?
 - b. The Cl atom in ClO_3^- ?

* * * * *

Redox Definitions

Redox reactions are a combination of *reduction* and *oxidation*.

1. **Oxidation** is the **loss** of electrons.
2. **Reduction** is the **gain** of electrons.

In a redox reaction, *electrons transfer* from one particle to another: one particle must be oxidized (lose its electrons), and another particle must be reduced (gain the transferred electrons).

The term *oxidation* is derived from the fact that molecular oxygen (O_2) is the agent most often used to remove electrons from substances. The term *reduction* reflects a chemical reaction important since prehistoric times. When metal ions found in ores are heated with substances that can donate electrons to the ions, the volume of the ore is seemingly *reduced*, as metals, valued for uses including ornaments and weaponry, are formed.

Oxidation Numbers

Oxidation numbers can be assigned to each individual atom in a chemical particle. Oxidation numbers can be thought of as the *charge* that each atom would have in a molecule or ion if each atom were a monatomic *ion*.

In reality, individual atoms most often do *not* behave as monatomic ions in chemical particles. However, assigning oxidation numbers can help in tracking the gain or loss of electrons that is the key to explaining many chemical reactions.

RULES for Assigning Oxidation Numbers

The goal is to assign an oxidation number (Ox#) to each *individual atom* in a substance.

1. In an *element*, each atom is assigned an oxidation number of *zero*.

Elements are neutral molecules with only one kind of atom. Examples of element formulas include C, Na, O₂, S₈, and Cl₂ (see Lessons 7B and 22D for additional information on elements). In those formulas, all atoms are assigned an oxidation number of zero.

2. For complex ionic compounds, write the formulas for each *ion* in the compound, then assign oxidation numbers to the atoms in the ions based on the rules below.
3. In a *monatomic ion*, the oxidation number of the atom is the *charge* on the ion.

Examples: In Na⁺, the Na atom is assigned an oxidation number of +1.

In Al³⁺, the Al is assigned an oxidation number of +3;

In S²⁻, the S is assigned an oxidation number of -2.

4. For particles that are *not* elements or monatomic ions, oxidation numbers are assigned as follows.
 - a. Each oxygen (O) atom is assigned a -2 (except in peroxides, where O is -1).
 - b. Each hydrogen (H) atom is assigned a +1, except in metallic hydrides (compounds of a metal atom and hydrogen) where H is -1.
 - c. Each alkali metal atom is assigned a +1, and each column 2 atom is assigned a +2.
 - d. All other oxidation numbers are selected to make the *sum* of the oxidation numbers equal the *overall charge* on the molecule or ion.

Example: What is the oxidation number for the Mn atom in MnO₄⁻?

In these lessons, we will use a labeling system where the oxidation number for the *each atom* is shown above the particle symbol, and the *total* for the atoms of that kind is shown below.

Each O has a -2 oxidation number. The total of the oxidation numbers for the oxygens is therefore -8.

Each:	-2
Formula for the particle:	MnO ₄ ⁻
Total:	-8

The total of the oxidation numbers must add up to the -1 charge on the ion. The oxidation number on the Mn atom must therefore be +7.

Each:	+7	-2
Formula for the particle:	MnO ₄ ⁻	
Total:	+7	-8 ← must add up to -1.

5. The oxidation number of *each atom* in a chemical particle is the total of the oxidation numbers for the atoms of that kind, *divided* by the number of atoms of that kind.

Example: In neutral N_2O_4 , each O is assigned a -2 oxidation number, and the total of the oxidation numbers for the oxygens is -8 .

Each:	+4	-2	
Formula for the particle:	N_2	O_4	
Total:	+8	-8	← must add up to 0.

The total of the oxidation numbers must add up to the zero charge on the molecule. The oxidation number on *each* N atom must therefore be **+4**.

Commit to memory the rules above and then, without looking back, apply the rules to this problem.

Q1. What is the oxidation number for the nitrogen atom in NO_2 ?

* * * * * (the * * * mean: write your answer, then check the answer below.)

Answer

- For any particle that is not an element or monatomic ion, first assign *individual*, then *total* oxidation numbers, for O, H, and column one and two atoms.
- Each oxygen atom in NO_2 has an oxidation number of -2 ; for the two oxygens, the total of the oxidation numbers is -4 .

Oxidation number of each atom:	-2	
Formula for the particle:	NO_2	
Total oxidation #'s for those atoms:	-4	(individual Ox#) x (subscript) = total)

To calculate the Ox# for the nitrogen atom, use the rule that the *total* numbers on the bottom must add up to the overall charge on the particle, in this case = zero.

The *total* of the oxidation numbers for all of the nitrogens must therefore be **+4**. Since we have only one nitrogen, its *individual* oxidation number must be **+4**.

Oxidation number of <i>each</i> atom:	+4	-2	
Formula for the particle:	NO_2		
Total oxidation numbers for those atoms:	+4	-4	← must total to charge on particle.

Try one more problem using the rules above.

Q2. Find the oxidation number of the nitrogen atom in ammonium ion, NH_4^+ .

* * * * *

Answer

- The overall charge on the NH_4^+ particle is **+1**; the total for the oxidation numbers must add up to **+1**.
- For any particle that is not an element or monatomic ion, assign *individual*, then *total* oxidation numbers, for O, H, and column one and two metal atoms.

In NH_4^+ , each hydrogen atom has an oxidation number of +1; the total oxidation numbers for the four hydrogens is +4.

Oxidation number of **each** atom: +1
 Formula for the particle: NH_4^+
Total Ox# for those atoms: +4 [(individual Ox#) x (subscript) = total]

- c. Calculate the *total*, then *individual* oxidation numbers for the remaining atoms.

Since the total charges for all the atoms (on the bottom) must add up to +1, the total for all nitrogens must be -3. Since there is only one nitrogen atom, its individual oxidation number must be -3.

Oxidation number of each atom: $\boxed{-3} + 1$
 Formula for the particle: NH_4^+
 Total oxidation numbers for those atoms: $-3 + 4 \leftarrow$ must add up to +1.

Summary: Assigning Oxidation Numbers

- Each atom in an element is assigned an oxidation number of zero.
- The oxidation number of an atom in a monatomic ion is the charge on the ion.
- To assign other oxidation numbers to individual atoms,
 - Write the *individual*, then *total* Ox# on O, H, and column *one* and *two* atoms;
 - Assign remaining *totals* for atoms that add up to the charge on the particle.
 - Write the oxidation numbers for each *individual* remaining atom.

Practice: Find the oxidation number of each individual atom for the atoms specified in these compounds. Do at least every other problem. Answers are on the following page.

- | | | | |
|---------------------------------|------------|--|------------|
| 1. CO_2 | C = _____ | 9. S_8 | S = _____ |
| 2. Br_2 | Br = _____ | 10. $\text{C}_2\text{O}_4^{2-}$ | C = _____ |
| 3. Mn^{2+} | Mn = _____ | 11. H_3AsO_4 | As = _____ |
| 4. CaBr_2 | Br = _____ | 12. KMnO_4 | Mn = _____ |
| 5. NO_3^- | N = _____ | 13. ClO_3^- | Cl = _____ |
| 6. Na_2CrO_4 | Cr = _____ | 14. Mn_2O_7 | Mn = _____ |
| 7. $\text{Cr}_2\text{O}_7^{2-}$ | Cr = _____ | 15. Hg_2^{2+} | Hg = _____ |
| 8. H_2O_2 | O = _____ | (H ₂ O ₂ is hydrogen peroxide) | |

ANSWERS

Pretest: 1. Oxidation is the loss of electrons. 2. Cr = +6 3. Cl = +5

Practice

1. **C = +4**

Do individual, then total Ox#, for O, H, and column 1&2. Here, the total for O is $2 \times (-2) = -4$.
Do total, then individual, for remaining atoms. C must be +4, since CO_2 has a zero charge.

Oxidation number of *each* atom: $\boxed{+4} - 2$

Formula for the particle: CO_2

Total oxidation numbers for those atoms: $+4 - 4 \leftarrow$ must total to zero.

2. **Br = Zero.** Br_2 has one kind of atom and is neutral: it is an element. In elements, each atom has an oxidation number of zero.

3. **Mn = +2.** The Ox# on a monatomic ion is equal to its charge.

4. **Br = -1**

If you recognize that CaBr_2 is ionic, write its monatomic ions. Each charge is the oxidation number.

OR -- do individual, then total Ox# for O, H, and col. 1&2. Ca is column 2: $1 \times (+2) = +2$.

Do total, then individual, for remaining atoms. Br must be -1 due to the zero charge on CaBr_2 .

Oxidation number of **each** atom: $+2 -1$

Formula for the particle: CaBr_2

Total oxidation numbers for those atoms: $+2 - 2 \leftarrow$ must total zero.

5. **N = +5** Do O first. Each O is -2. $3 \times -2 = -6$. N must be +5 to equal the -1 total charge.

Each atom: $+5 -2$

Formula for the particle: NO_3^-

Total for those atoms: $+5 - 6 \leftarrow$ must total -1.

6. **Cr = +6** Each Na is +1 (Na is an alkali metal). Each O is -2.

Each atom: $+1 +6 -2$

Formula for the particle: Na_2CrO_4

Total for those atoms: $+2 +6 - 8 \leftarrow$ must total zero.

7. **Cr = +6** Each O is -2: $7 \times -2 = -14$. The 2 Cr must total +12, so **each** is +6.

Each: $+6 -2$

Formula for the particle: $\text{Cr}_2\text{O}_7^{2-}$

Total: $+12 -14 \leftarrow$ must total -2.

8. **O = -1.** In *peroxides*, oxygen has an Ox# of -1.

9. **S = Zero.** S_8 is an element: one kind of atom and neutral. Atoms in elements have Ox# of zero.

10. **C = +3** Each O is -2 : $4 \times -2 = -8$. The 2 Carbons must total $+6$, so **each** is $+3$.

Each:	+3	-2	
Formula for the particle:	C_2	O_4	$2-$
Total:	$+6$	-8	← must total -2.

11. **As = +5** Each H is $+1$. Each O is -2 . (Individual Ox#) times (subscript) = Ox# Total

Each:	$+1$	+5	-2
Formula for the particle:	H_3	As	O_4
Total:	$+3$	$+5$	-8 ← must total zero.

12. **Mn = +7** 13. **Cl = +5** 14. **Mn = +7** 15. **Hg = +1**

* * * * *

Lesson 15B: Balancing Charge

Prerequisites: Make sure that you have mastered *Lesson 10B: Balancing Equations*.

Pretest: On the remaining lessons in this module, if you think you know the topic, try the last few problems in the lesson. If you can do those, you may skip the lesson.

* * * * *

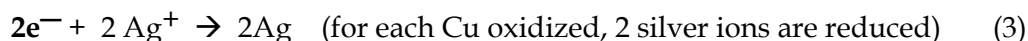
Half Reactions

In balanced redox equations, the number of electrons lost by one reactant particle must equal the number gained by another reactant particle. However, this key balanced transfer of electrons is hidden when a redox equation is written. To understand redox reactions, it is often necessary to separate the reaction into two parts in order to see and balance the number of electrons lost and gained.

For example, the oxidation of copper by silver ion can be represented as



This reaction is best understood as a result of two *half*-reactions. Each reactant copper atom loses two electrons, and each reactant silver ion gains one electron.



A redox **half-reaction** is a balanced equation that includes free electrons.

Every redox reaction is a result of *two* half reactions: one that loses electrons, and the other that gains the same number of electrons. Half-reactions break a redox reaction into its two components: the oxidation and the reduction.

In all chemical reactions, both atoms and charge must be conserved: in half-reactions and in complete redox reactions, the number and kind of *atoms*, and the total of the *charges*, must be the same on each side. Check each of the three reactions above to see that *both* atoms and total charge are balanced.

Balancing Charge In Half-Reactions

To balance redox reactions, it is often necessary to balance the half-reactions first.

To balance *half-reactions*,

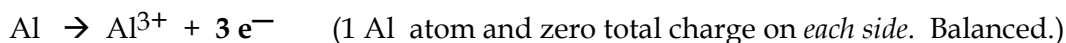
- first *add coefficients* to balance the atoms,
- then *add electrons* (e^-) to balance charge.

Apply those rules to this problem.

Q. Balance this half-reaction: $\text{Al} \rightarrow \text{Al}^{3+}$

★ ★ ★ ★ ★

A. Each side has one Al atom, so atoms are already balanced. The left side is neutral, and the right has a charge of 3+. For charge to be the same on each side, *you* must add three negative electrons to the right.



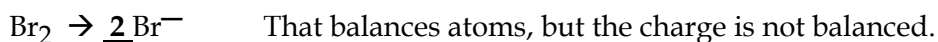
The number of *atoms* must be the same on each side of a balanced equation, but the number of *electrons* does *not*. Electrons are not atoms.

The total *charge* must be the same on each side of a equation. The electron in a half-reaction equation is a -1 charge that serves to balance charge. Try this problem.

Q. Balance this half-reaction: $\text{Br}_2 \rightarrow \text{Br}^-$

★ ★ ★ ★ ★

A. First balance *atoms* by adding *coefficients* (you cannot change the particle formulas). A coefficient of 2 must be added on the right.



The right side has a total charge of 2−, and the left is neutral. To balance charge in a half-reaction, two electrons must be added to the left.

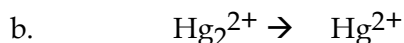


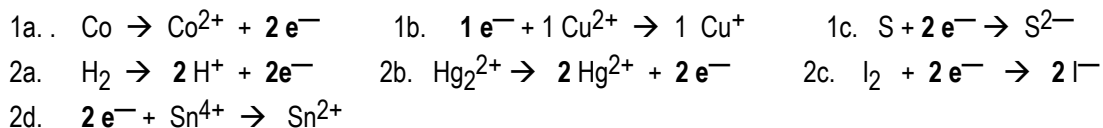
Practice: Check answers as you go.

1. Add electrons to balance these half-reactions.



2. Balance these half-reactions.



ANSWERS

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Lesson 15C: Oxidizing and Reducing Agents

Timing: Complete this lesson when you are asked to identify oxidizing and reducing agents. In some courses, this is done as part of a unit on ions and solutions. Other courses cover agents during later units on electrochemistry.

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Redox Agent Terminology

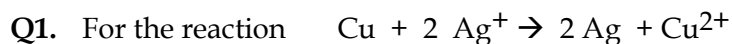
Redox reactions are a combination of **reduction** and **oxidation**. One reactant must be oxidized (*lose* its electrons), and another reactant **must** be reduced (*gain* the transferred electrons).

In a redox reaction, the reactant particle that is **oxidized** is the one that contains an atom that *loses* electrons. This particle is termed the **reducing agent**, because, by giving its electrons to another particle, it acts as an agent causing the other particle to be reduced.

The reactant particle that is **reduced** is the one that contains an atom that *gains* electrons. This particle is termed the **oxidizing agent**, because, by taking electrons from another particle, it is the agent that causes the other particle to be oxidized. On each side of a redox reaction, there must be one oxidizing agent and one reducing agent.

The following definitions must be memorized.

- **Oxidation** is the loss of electrons. **Reduction** is the gain of electrons.
- **Oxidizing agents** are particles that contain an atom that *accepts* electrons in a reaction; it removes electrons from another atom.
- **Reducing agents** are particles that contain an atom that *loses* electrons in a reaction by donating its electrons to another atom.
- In redox reactions, **reducing agents are oxidized**, losing their electrons, and **oxidizing agents are reduced**, gaining electrons.
- After learning these definitions, write your answers to the following questions.



1. Is the neutral copper being oxidized or reduced?

* * * * *

A. In the reaction, Cu becomes Cu^{2+} ; it must lose 2 electrons to do so. As it loses electrons, the neutral Cu metal is being oxidized.

2. Is the Ag^{+} being oxidized or reduced?

* * * * *

- A. Since the neutral Cu reactant is being oxidized, the Ag^+ reactant must be being reduced.

The Ag^+ becomes Ag in the reaction. To do so, must gain an electron. Particles that gain electrons in a reaction are being reduced.

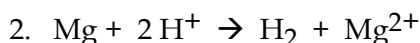
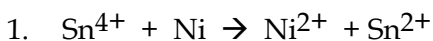
3. Is the Ag^+ acting as an oxidizing agent or a reducing agent?

* * * * *

- A. Since Ag^+ is being reduced, it is an oxidizing agent. By accepting the electron from neutral copper, Ag^+ causes the Cu to be oxidized.

Practice A

Add a label below each *reactant* identifying it as an oxidizing agent (OA) or reducing agent (RA), then circle the reactant particle being oxidized.

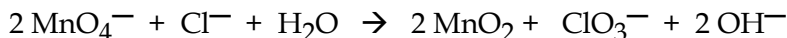


Identifying Agents Using Oxidation Numbers

In more complex reactions, it is often helpful to assign oxidation numbers to atoms to identify which particles are being oxidized and reduced.

Q2. For this redox reaction,

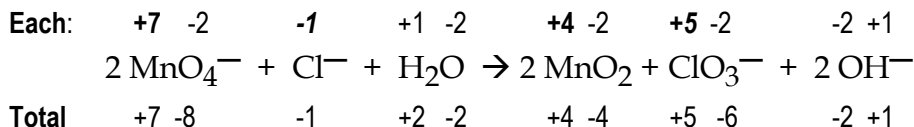
- assign oxidation numbers to each atom,
- write the formula for the reactant particle that is the oxidizing agent, and
- write the formula for the reactant being oxidized.



* * * * *

Answer

- a. Assigning oxidation numbers,



- The Mn atom is +7 left and +4 right, so it gains 3 electrons. MnO_4^- contains the atom that is being reduced, so MnO_4^- ion is an **oxidizing agent**.
- The Cl atom is -1 in the reactants and +5 in the products. The Cl^- ion is therefore losing 6 electrons in the reaction and is the reactant being **oxidized**.

In general, strong oxidizing agents include O_2 , CrO_4^{2-} , $\text{Cr}_2\text{O}_7^{2-}$, and substances or ions that include *per-* as a prefix in their names indicating a high number of oxygens in the particle, such as permanganate. Elements that are metals are often used as reducing

agents. However, to determine the role of particles in a reaction, rather than memorize a list, it is best to assign oxidation numbers and analyze the gain and loss of electrons.

* * * * *

Flashcards

Make needed cards from the list below. Run the new cards 3 days in a row, and again before each quiz or test on this material.

One-way cards (with notch)

Back Side -- Answers

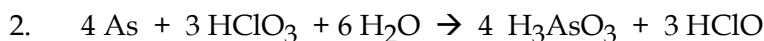
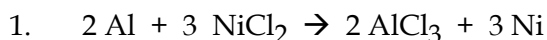
Oxidation number for atoms in an element	Zero
Normal oxidation number for oxygen atoms	-2
Oxidation number for O atoms in peroxides	-1
Normal oxidation number for hydrogen atoms	+1
Oxidation number for H atoms in hydrides	-1
Oxidation # for alkali metals in compounds	+1
Oxidation # for column 2 atoms in compounds	+2
The oxidation number of a monatomic ion is	Its charge
In reactions, reducing agents are	Oxidized
In reactions, oxidizing agents are	Reduced
Total oxidation numbers must add up to	The charge on the particle
The fundamental law of redox balancing	Electrons lost must equal electrons gained

Two-way cards (with *out* notch):

Reduction	The gain of electrons
Oxidation	The loss of electrons
Oxidizing agent	Particle that accepts electrons
Reducing Agent	Particle that gives away electrons

Practice B

Assign oxidation numbers, label each reactant as an oxidizing agent (OA) or reducing agent (RA), and circle the reactant being oxidized.



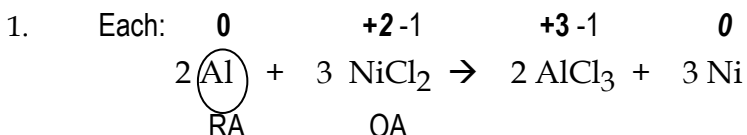
ANSWERS**Practice A**

In **1a**, Sn^{4+} changes to Sn^{2+} , so it must gain 2 electrons and is being reduced. Because Sn^{4+} is accepting electrons, it is the oxidizing *agent*.

The Ni loses 2 electrons in reacting, so it is being *oxidized*. Because Ni gives electrons to the Sn^{4+} , it is a reducing agent. The particle being oxidized is the reducing agent.

In **1b**, the Mg on the left loses two electrons as it forms Mg^{2+} . It is therefore being *oxidized*. Because it is giving away its electrons, it is the reducing agent.

That means the H^+ must be the oxidizing agent. Two H^+ ions gain 2 electrons to form neutral hydrogen. In a redox reaction, the reactants must contain *both* an RA and an OA.

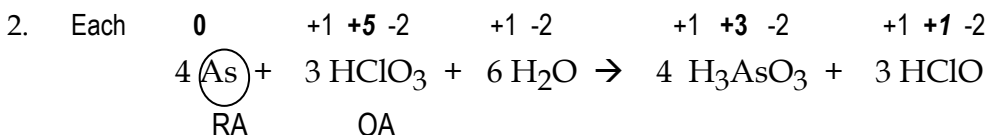
Practice B

Neutral Al and Ni atoms are elements. Atoms in elements have a zero oxidation number.

NiCl_2 and AlCl_3 are both *ionic* compounds because they combine metal and non-metal atoms. All of their atoms are monatomic ions. The oxidation numbers for the atoms are the charges of the ions.

Since the **Al** goes from neutral to 3+, it loses 3 electrons. A particle that loses electrons is being oxidized and is acting as a reducing agent.

The other reactant (NiCl_2) must therefore be oxidizing agent. The Ni^{2+} ion gains two electrons.



The **As** atom is an element with a zero oxidation number. After reaction, it has a +3 oxidation number. The As loses 3 electrons, is oxidized, and acts as a reducing agent.

The **Cl** atom goes from a +5 oxidation number in the reactants to +1 in the products. It gains 4 electrons. Particles that contain atoms that gain electrons in a reaction are reduced. Because they accept electrons, those particles are oxidizing agents.

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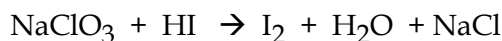
Lesson 15D: Balancing Redox Using Oxidation Numbers

Trial and error can balance all chemical equations, but for many redox reactions, trial and error can be a slow process. Oxidation numbers can speed balancing by suggesting coefficients. A key principle in oxidation-reduction reactions is

The number of electrons lost by the reducing agent must equal the number gained by the oxidizing agent.

For unbalanced redox equations, by assigning oxidation numbers, the gain and loss of electrons among atoms can be seen. Coefficients can then be assigned that balance the electron gain and loss, and balancing can be completed by trial and error.

The oxidation-number method of balancing is best learned by example. Using the following redox equation,

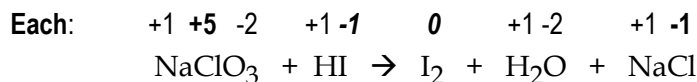


complete the following steps in your notebook.

Using Oxidation Numbers to Balance Redox Reactions

1. Above the atoms, write the oxidation numbers for *each atom* in the equation.

* * * * *



2. Identify the *two* atoms on *each side* that *change* their oxidation number in going from reactants to products.

* * * * *

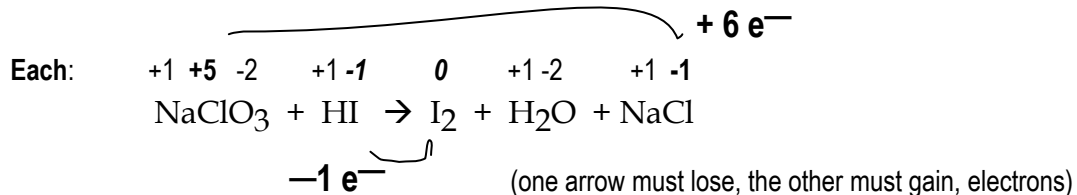
The two atoms that change oxidation numbers are **I** and **Cl**.

3. For each of those two atoms, draw arrows connecting the atom on one side to the same atom on the other side. Draw one arrow above and the other below the equation. Label each arrow with the *electron change* that must take place in going from one oxidation number to the other.

* * * * *

To go from +5 to -1, each chlorine must gain six negative electrons.

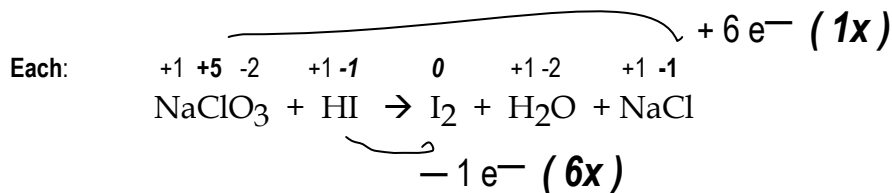
To go from -1 to zero, each iodine atom must lose one electron.



4. Calculate the multiplier, based on the lowest common multiplier (LCM), by which each loss or gain should be multiplied so that the *total* electrons lost and gained are *equal*.

* * * * *

Multiply the I arrow by 6 and the Cl arrow by 1.



5. Re-write the original equation, using the 2 arrow multipliers as **trial coefficients** for the two particles connected by the arrow.

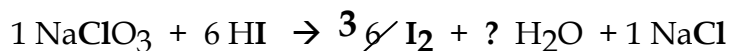
* * * * *



6. The **subscript tweak**. For each *atom* that *changed* oxidation number, if its subscript is *not* 1, divide the trial coefficient in front of its *particle* by the subscript of that atom.

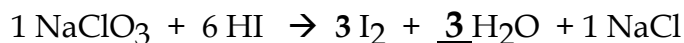
* * * * *

In this problem, we have a subscript tweak in one case out of 4. One of the I atoms has a subscript of two, so divide the trial coefficient of its particle by 2.



7. Add coefficients to finish the balancing by trial and error.

* * * * *



8. **Check:** 1 Na, 1 Cl, 3 O, 6 H, 6 I atoms on each side. Neutral charge on each side. The equation is balanced. Done!

* * * * *

Oxidation number balancing gives a few *trial* coefficients. These trial coefficients are usually good hints, but they are not final answers.

In *all* methods of balancing, it is necessary to modify trial coefficients if needed, by trial and error, until the atoms and overall charge are the same on each side.

This method will become easier with practice.

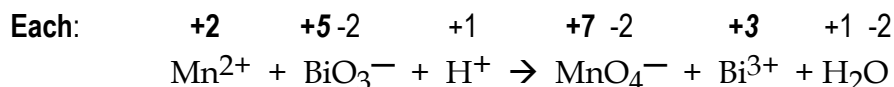
Practice: Write a “key word” summary of the steps above. Learn it. Then write your summary from memory and use it to balance these redox equations. If you need help, peek at the answer and try again. Save one problem for your next practice session.

- $\text{Mn}^{2+} + \text{BiO}_3^- + \text{H}^+ \rightarrow \text{MnO}_4^- + \text{Bi}^{3+} + \text{H}_2\text{O}$
- $\text{KI} + \text{O}_2 + \text{H}_2\text{O} \rightarrow \text{I}_2 + \text{KOH}$
- $\text{Co} + \text{H}^+ + \text{NO}_3^- \rightarrow \text{Co}^{2+} + \text{N}_2 + \text{H}_2\text{O}$
- $\text{Zn} + \text{H}^+ + \text{MnO}_4^- \rightarrow \text{Zn}^{2+} + \text{Mn}^{2+} + \text{H}_2\text{O}$
- $\text{Al} + \text{HCl} \rightarrow \text{Al}^{3+} + \text{Cl}^- + \text{H}_2$
- Write the formula for the reactant that is the reducing agent in each of your balanced equations.
- Write the formula for the reactant that is being oxidized in each of your balanced equations.

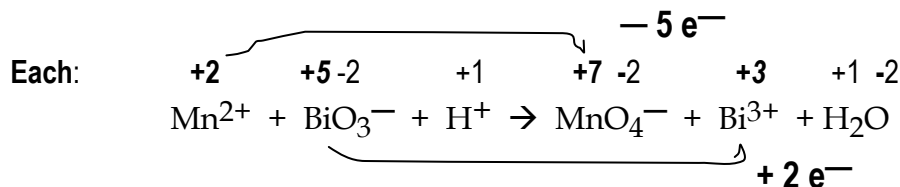
ANSWERS

Problem 1

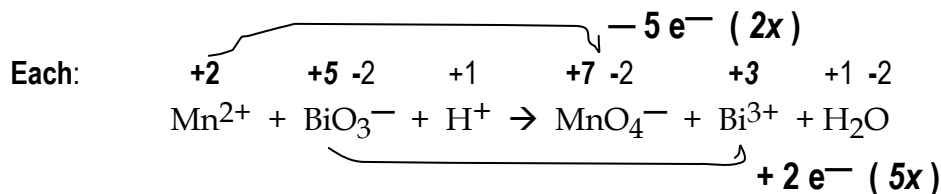
- Calculate the oxidation numbers for each atom in the equation.



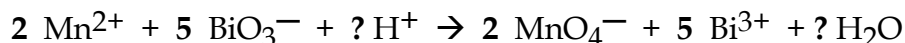
- Identify two atoms that change their oxidation numbers. **Mn and Bi.**
- Add arrows connecting each atom that changed. Label with the *electron change*.



- Add the lowest arrow multipliers that result in the electron loss and gain being equal.

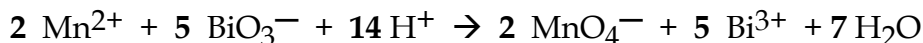


- Re-write the original equation. Use each arrow multiplier as a *coefficient* for the *particles* connected by the arrows.



- The *subscript tweak*. Not needed – none of the Bi or Mn atoms have subscripts.
- Complete balancing by trial and error.

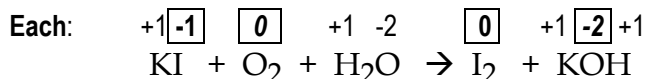
15 O on the left. Must be 7 H₂O on right to get 15 O on right. 7 H₂O means 14 H⁺ on left.



- Check:** 2 Mn, 5 Bi, 15 O, 14 H atoms on each side. +13 charge on each side. Balanced.

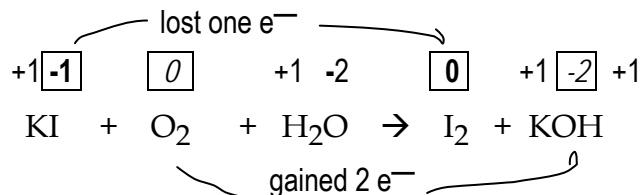
Problem 2

- Calculate the Ox# for each atom.

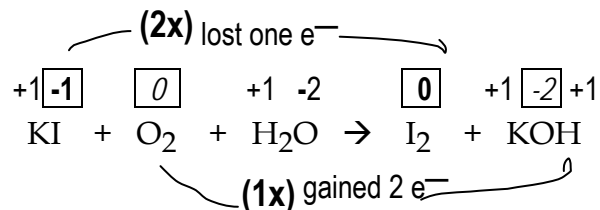


O₂ and I₂ are *elements*: neutral particles containing only one kind of atom.

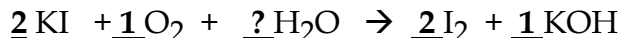
- Identify two atoms on each side that change their oxidation number. **I and O.**
- Write the electron change for each atom.



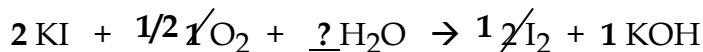
- Calculate an LCM multiplier for each to get the same number of electrons in both.



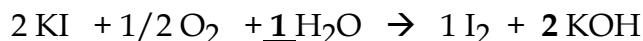
- In the original equation, add the multipliers as coefficients for connected particles.



- The *subscript tweak*. Since both the O on the left and the I on the right have a subscript of 2, divide the coefficient of their particles by 2. Fractions are OK at this point.



- Finish by trial and error. With 2 K on the left, there must be 2 on the right, so adjust the trial KOH coefficient. With 2 H now on the right, one water is needed on the left.



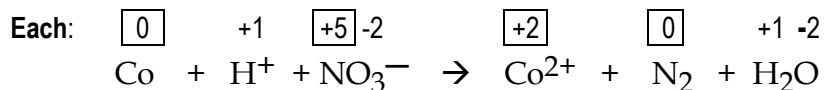
- Check:** 2 K, 2 I, 2 O and 2 H atoms on each side. Zero charge on each side. Balanced!

If desired, you may multiply all coefficients by 2 to eliminate the fraction. Fractions are permitted, but calculations based on balanced equations are easier without the fractions.

Note that the trial coefficients had to be adjusted at the end. The coefficients found by oxidation numbers are good *hints*; but always adjust by trial and error at the end, if needed, until the equation is balanced.

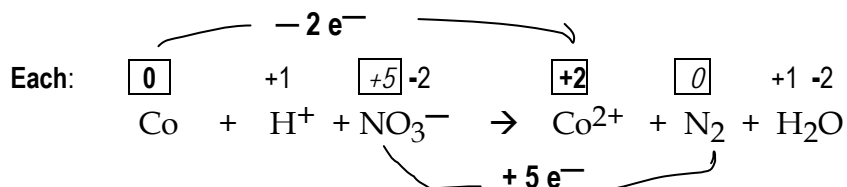
Problem 3

1. Calculate the Ox# for each atom.

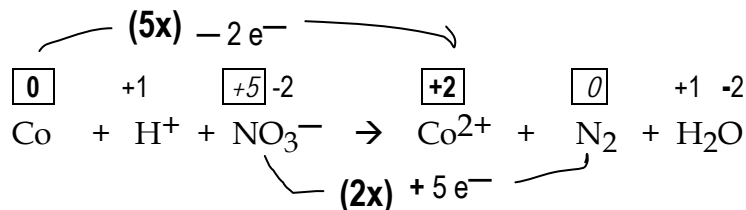


2. Identify two atoms that change oxidation number.
- Co**
- and
- N**

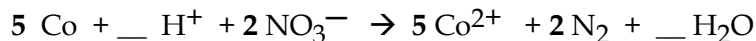
3. Write the electron change for each atom.



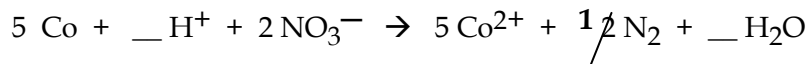
4. Add the lowest common arrow multiplier that results in electron loss = gain.



5. To the original equation, add the multipliers as
- coefficients*
- for both connected particles.

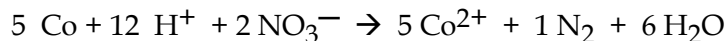


6. The
- subscript tweak*
- . Since N on the right has a subscript 2, divide its coefficient by 2.

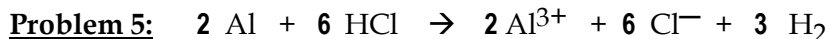
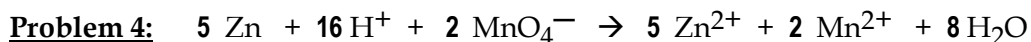


7. Complete by trial and error.

The 6 O atoms on the left require 6 H₂O on the right. That means 12 H⁺ must be on the left.



- 8.
- Check:**
- 5 Co, 12 H, 2 N and 6 O atoms on each side.
- +10**
- charge on each side.
- Balanced!**



Problem 6: The reducing agent is the reactant particle that is losing electrons and donating them to another particle.

- 1.
- Mn²⁺**
- is losing 5 electrons. 2.
- I⁻**
- is losing 1 electron. 3.
- Co**
- 4.
- Zn**
- 5.
- Al**

Problem 7: The particle being oxidized is the reducing agent. Answers are the same as Problem 6.

* * * * *

Lesson 15E: Redox Stoichiometry

Timing: Do this module *if* you are assigned reaction amount calculations (stoichiometry) involving redox reactions. If your current assignments include learning to balance redox using *half*-reactions, first complete Module 16, then return to this lesson.

* * * * *

Redox reactions are more difficult to balance than most chemical equations. However, once a redox equation is balanced, calculations for redox reactions can be done using the same steps as other stoichiometry.

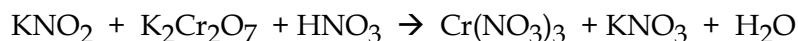
Stoichiometry Steps

1. If the moles for two reactant particles are supplied in the data without a limiting reactant identified, solve using a *rice* table. However, most *redox* reaction calculations are reactions to a stoichiometric equivalence point (endpoint). In such cases, *rice* tables work (as always), but conversion stoichiometry solves more quickly.
2. If you want a *single* unit, solve by the 7 *single*-unit stoichiometry steps (Lesson 10E):
WANTED and DATA, balance and bridge, convert units to *moles* to *moles* to *units*.
3. If you want a *ratio* unit, solve for the top and bottom WANTED units separately, then divide (Lesson 12D).
4. For reactions involving *ions*, write the balanced equation using *solid* and then *separated* formulas for the reactants and products.
5. For some stoichiometry, you may not need to balance the entire equation. You will only need to balance enough to find the *coefficients* for the WANTED and *given* substances that are used in the *bridge* conversion.

Practice: If you get stuck, read a bit of the answer, then try again.

1. In the problems below, for *which* reactants can you find the moles initially present based on the data supplied (without using a stoichiometry mole-to-mole conversion),
 - a. In problem 2?
 - b. In problem 3?
2. In a redox titration of 40.0 mL of FeCl₂ solution, the endpoint is reached when 24.4 mL of 0.200 M KMnO₄ is added. The unbalanced equation is

$$\text{KMnO}_4 + \text{FeCl}_2 + \text{HCl} \rightarrow \text{FeCl}_3 + \text{MnCl}_2 + \text{H}_2\text{O} + \text{KCl}$$
 - a. What is the original [FeCl₂]?
 - b. Which reactant is the oxidizing agent?
 - c. Which reactant is being oxidized?
3. How many mL of acidic 0.200 M K₂Cr₂O₇ solution are needed to titrate a sample of 0.851 grams of KNO₂? The unbalanced equation is



d. (Solve for the *final*/WANTED unit using the two amounts found for the sample.)

$$\text{WANT: } ? \frac{\text{mol FeCl}_2}{\text{L FeCl}_2 \text{ soln.}} = \frac{2.440 \times 10^{-2} \text{ mol FeCl}_2}{40.0 \times 10^{-3} \text{ L FeCl}_2} = \boxed{0.610 \text{ M FeCl}_2}$$

2b. The oxidizing agent is the reactant particle (on left side) that is gaining electrons in the reaction. Since MnO_4^- is the particle with an atom that is gaining e^- (Mn), the reactant gaining e^- is the KMnO_4 .

2c. The balancing shows that Fe^{2+} is the ion losing an electron, so FeCl_2 is the reactant being oxidized.

3a. Use the rule: when *titration* or an *endpoint* or *equivalence point* is involved, use conversion stoichiometry to solve, or the rule: since the moles are supplied for only one reactant, use conversion stoichiometry.

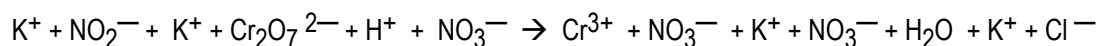
WANT: ? mL $\text{K}_2\text{Cr}_2\text{O}_7$ soln.

DATA: 0.851 g KNO_2 (your single-unit *given*)

0.200 mol $\text{K}_2\text{Cr}_2\text{O}_7 = 1 \text{ L } \text{K}_2\text{Cr}_2\text{O}_7$ soln.

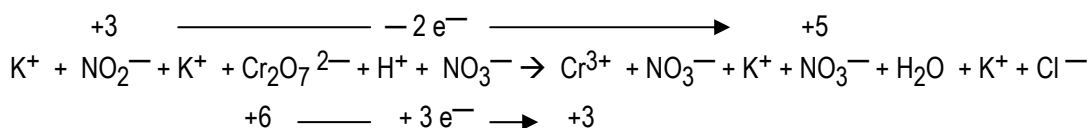
85.1 g $\text{KNO}_2 = 1 \text{ mol } \text{KNO}_2$ (g prompt)

Balance: First convert the solid formulas to the ions they separate into in solution.

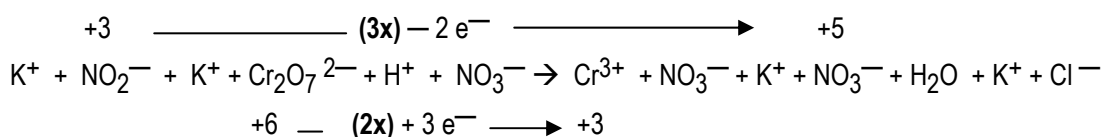


Solve using oxidation numbers as below, or by using half-reactions.

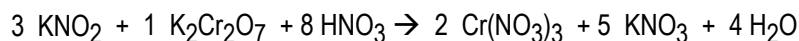
Assign oxidation numbers; identify the two atoms that change oxidation numbers.



N and Cr. Decide the multipliers that equalize the electrons lost and gained .



Use the multipliers as trial coefficients and finish balancing by trial and error.



Note subscript tweak on the coefficient of $\text{K}_2\text{Cr}_2\text{O}_7$. Check: 5 K, 11 N, 37 O, 2 Cr, 8 H on each side.

Bridge: **1 mol $\text{K}_2\text{Cr}_2\text{O}_7 = 3 \text{ mol } \text{KNO}_2$**

SOLVE: (Since a single unit is wanted, solve with 7-step single-unit stoichiometry, chaining conversions.)

$$\begin{aligned} ? \text{ mL } \text{K}_2\text{Cr}_2\text{O}_7 &= 0.851 \text{ g } \text{KNO}_2 \cdot \frac{1 \text{ mol } \text{KNO}_2}{85.1 \text{ g } \text{KNO}_2} \cdot \frac{1 \text{ mol } \text{K}_2\text{Cr}_2\text{O}_7}{3 \text{ mol } \text{KNO}_2} \cdot \frac{1 \text{ L } \text{K}_2\text{Cr}_2\text{O}_7}{0.200 \text{ mol } \text{K}_2\text{Cr}_2\text{O}_7} \cdot \frac{1 \text{ mL}}{10^{-3} \text{ L}} = \\ &= \mathbf{16.7 \text{ mL } \text{K}_2\text{Cr}_2\text{O}_7 \text{ soln.}} \end{aligned}$$

* * * * *

Summary: Redox Balancing and Oxidation Numbers

1. Redox definitions:
 - **Oxidation** is the loss of electrons. **Reduction** is the gain of electrons.
 - **Oxidizing agents** are particles that contain an atom that *accepts* electrons in a reaction; it removes electrons from another atom.
 - **Reducing agents** are particles that contain an atom that loses electrons in a reaction by donating its electrons to another atom.
 - In redox reactions, **reducing agents** are **oxidized**, losing their electrons, and **oxidizing agents** are **reduced**, gaining electrons.
2. Rules for assigning oxidation numbers.
 - a. Each atom in an element is assigned an oxidation number of zero.
 - b. The oxidation number of an atom in a monatomic ion is the charge on the ion.
 - c. For particles that are not elements or monatomic ions, oxidation numbers are assigned as follows:
 - i. Each oxygen (O) atom is assigned a -2 (except in peroxides, O is -1).
 - ii. Each hydrogen (H) atom is assigned a $+1$, except in metallic hydrides (compounds of a metal atom and hydrogen), where H is -1 .
 - iii. Each alkali metal atom is assigned a $+1$. Each column 2 atom is assigned a $+2$.
 - iv. All other oxidation numbers are selected to make the sum of the oxidation numbers equal the overall charge on the molecule or ion.
 - d. The oxidation number of each individual atom in a chemical particle is the total of the oxidation numbers for the atoms of that kind, divided by the number of atoms of that kind.
3. A key principle in oxidation-reduction reactions is

The number of electrons lost by the reducing agent must equal the number gained by the oxidizing agent.

To balance redox reactions, coefficients must be added so that the number of electrons lost by the reducing agent equals the number gained by the oxidizing agent.

4. In balancing redox reactions, assigning oxidation numbers to assist in balancing provides *trial* coefficients. Always adjust the coefficients by trial and error at the last step in order to complete the balancing.

#

Module 16: Half-Reaction Balancing

Timing: Module 16 covers redox balancing using *half-reactions*. Some courses assign this topic after oxidation numbers. Others assign half-reactions as part of electrochemistry later in the course. Do this unit when *half-reactions* are assigned in *your* course.

Prerequisite: If you cannot solve these problems easily, review Lessons 15A, 15B, and 15D before starting Module 16. Answers are at the end of Lesson 16A.

- Define reduction.
- Balance this half-reaction: $F_2 \rightarrow F^-$
- Write the oxidation number of
 - Each Cl atom in Cl_2O_5
 - The Mn atom in MnO_4^-
- Label each *reactant* as an oxidizing agent (OA) or reducing agent (RA). Circle the reactant being oxidized.
 - $Sn^{4+} + Co \rightarrow Co^{2+} + Sn^{2+}$
 - $Ca + 2 H^+ \rightarrow H_2 + Ca^{2+}$

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Lesson 16A: Balancing Half-Reactions: The CA-WHe! Method

Balancing Half-Reactions That Include Acid and Water

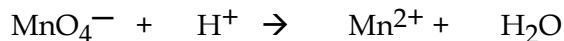
Many redox reactions occur in aqueous solutions, and water is often a term in the reactants or products. Many redox reactions carried out in aqueous solutions need acidic conditions for the desired reaction to take place. In these cases, both the redox half-reactions and the overall reaction may include H^+ ions and H_2O in the reactants and/or products.

The steps for balancing half-reactions which include H^+ and H_2O are the same as for balancing other half-reactions.

- First add *coefficients* to balance *atoms*,
- then add *electrons* to balance *charge*.

Using those two steps, solve the following problem.

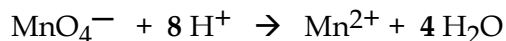
Q. Balance this half-reaction.



* * * * * (the * * * means: cover below, write *your* answer, *then* check below.)

Each side has one Mn atom: Mn is balanced. The left has 4 oxygens, so the right must have a coefficient of 4 for water.

That gives 8 H atoms on the right, so the H^+ coefficient on the left must be 8 for the H atoms to balance. The atoms are now balanced:



but the charge is *not* balanced. Add electrons to the equation to balance the charges.

* * * * *

Since the left charges total 7+ and the right 2+, we must add 5 electrons to the left in order for the charges to balance.

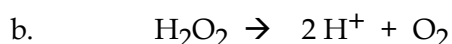
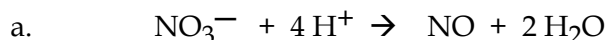


The net charge is now +2 on each side. Atoms *and* charge are now balanced, so the half-reaction is balanced.

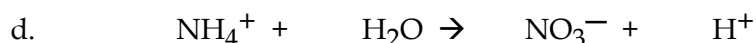
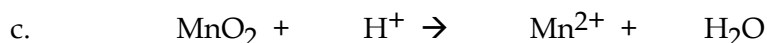
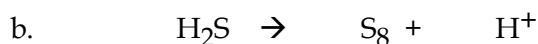
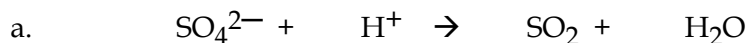
Practice A

Do every other problem. Check answers as you go. Need more practice? Do more.

1. Add electrons to balance these half-reactions.



2. Balance these half-reactions.



Balancing Half-Reactions By the CA-WHe Method

In textbook and test problems, for redox reactions run in aqueous solutions, the H^{+} , H_2O , and electrons needed to balance a half-reaction often are not supplied. In those cases, you will need to add those terms to balance the half-reactions.

You can construct half-reactions by using

The CA-WHe! Method

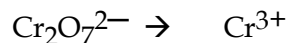
To balance redox half-reactions,

1. first balance the "central atom" (CA), usually one that is **not** O or H. Then,
2. add **W**ater if needed to balance the **oxygens**.
3. Add **H**⁺ if needed to balance the **hydrogens**.
4. Add **e**lectrons to balance the **charge**.
5. *Check* that atoms and charge are the same on each side.

It helps to memorize: "To balance half-reactions, balance the central atom, then WHe!"

Use the CA-WHe! method on the following example.

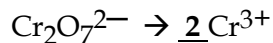
Q. Balance this half-reaction.



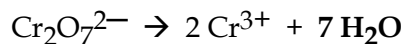
* * * * *

Answer

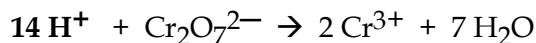
1. First, balance the central atom (usually an atom that is not O or H).



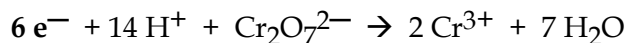
2. Add Water to balance *oxygen* atoms.



3. Add H⁺ (protons) to balance *hydrogen*.



4. Add e⁻ to balance *charge*.

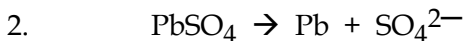


5. After balancing, check to see that the count for each kind of *atom* is the same on each side and the total *charge* is the same on each side.

Check: 14 H, 2 Cr, 7 O on each side; +6 charge on each side. Balanced.

Practice B

Complete the balancing of these half-reactions, adding H⁺ and H₂O if needed. Assume a particle is neutral if no charge is shown. Do the odd problems, checking your answers as you go. Do the evens during your next practice session.



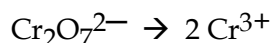
Balancing Half-Reactions That Include Hydroxide Ion (OH⁻)

Some redox reactions require *basic* conditions for the desired reaction to take place. In these cases, the redox half-reactions may include OH⁻ ions instead H⁺ ions. The steps to balance these half-reactions are

1. First balance by the CA-WHe! method using H⁺ and H₂O.
2. Then *neutralize* the H⁺: using the H⁺ coefficient, add that number of OH⁻ ions to *each side*. This “neutralization of H⁺” will replace the number of H⁺ ions in the reaction with an equal number of H₂O molecules, and add the same number of OH⁻ ions to the side opposite the original H⁺.
3. Adjust the H₂O coefficients on each side if needed.

Use those steps to do the following problem. If you get stuck, read part of the answer below, then complete the problem.

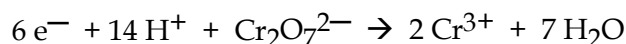
Q. Balance this half-reaction using OH⁻ ions instead H⁺.



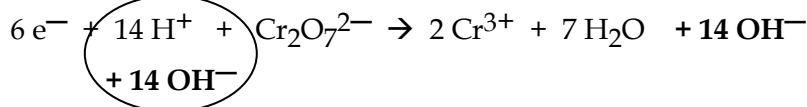
* * * * *

Answer

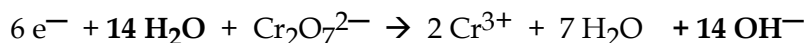
1. Balance the half-reaction using the CA-WHe! method.



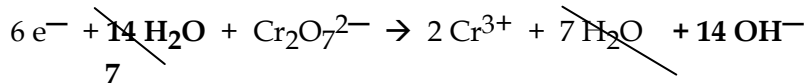
2. Neutralize the H⁺ by adding OH⁻ to each side.



3. React the H⁺ and OH⁻ to form H₂O.



4. Adjust the count of the H₂O. 7 H₂O cancel on each side.



5. Check: 14 H, 2 Cr, 14 O, 8 negative charges on each side.

When to Use OH⁻ Ions to Balance

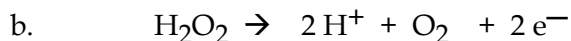
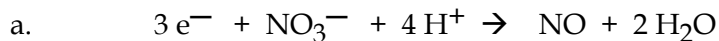
In problems balancing redox reactions, if the problem says

- “In acid solution,” use H⁺ ions to balance the half-reactions.
- “In base solution” or involves OH⁻ ions in the products or reactants, use OH⁻ ions to balance the half-reactions.

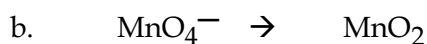
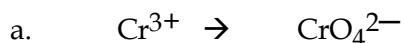
If the problem does not specify an acidic or basic solution, assume an acidic solution.

Practice C: Do every other problem. Save the rest for your next practice session.

1. Modify these half-reactions to balance using OH^- ions instead H^+ .



2. Balance these half-reactions assuming basic solutions.



ANSWERS

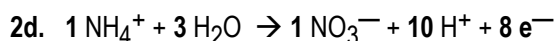
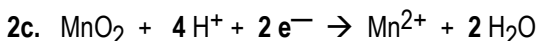
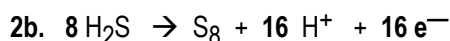
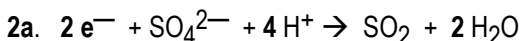
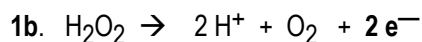
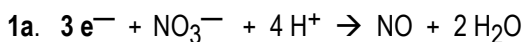
Pretest

1. The gain of electrons. 2. $2\text{e}^- + \text{F}_2 \rightarrow 2\text{F}^-$ 3. Cl = +5, Mn = +7



Practice A

The total of the charges on each side must balance.



Practice B

1. First balance the Central Atom: $2 \text{ClO}_3^- \rightarrow \text{Cl}_2$

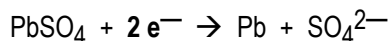
Then WHe: Add water to balance O: $2 \text{ClO}_3^- \rightarrow \text{Cl}_2 + 6 \text{H}_2\text{O}$

H⁺ to balance H: $2 \text{ClO}_3^- + 12 \text{H}^+ \rightarrow \text{Cl}_2 + 6 \text{H}_2\text{O}$

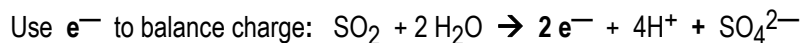
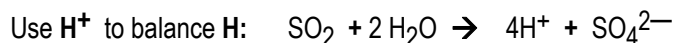
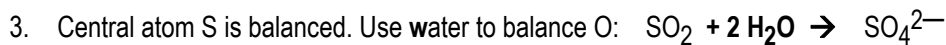
e⁻ to balance charge: $2 \text{ClO}_3^- + 12 \text{H}^+ + 10 \text{e}^- \rightarrow \text{Cl}_2 + 6 \text{H}_2\text{O}$

Check: 2 Cl, 6 O, 12 H, zero charge on each side. Balanced.

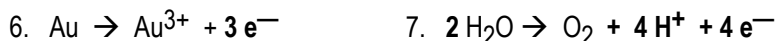
2. Central atom, O, and H are already balanced. Use e^- to balance charge.



Check: 1 Pb, 1 S, 4 O, 2 negative charges on each side. Balanced.

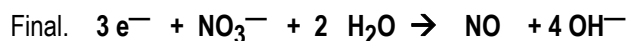
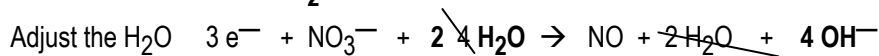
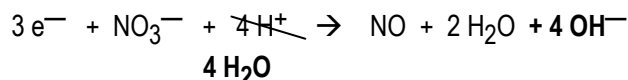


Check: 1 S, 4 O, 4 H, zero charge on each side. Balanced.

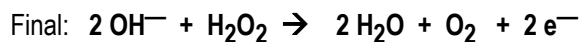
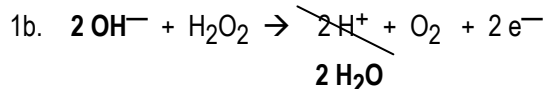


Practice C

1a. Add enough OH^- to each side to neutralize H^+ .



Check: 1 N, 5 O, 4H, 4 negative charges on each side. Balanced.

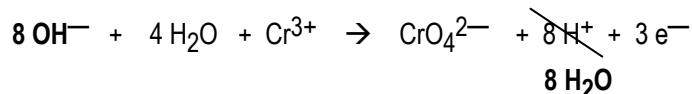


Check: 4 O, 4 H, 2 negative charges on each side. Balanced.

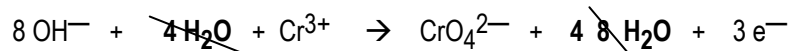
2a. "In basic solutions" means use OH^- ions. First balance using CA-WHe! method.



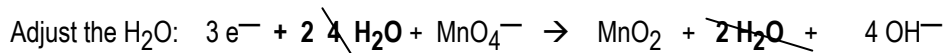
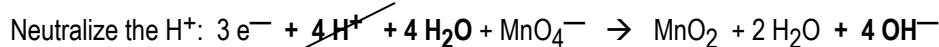
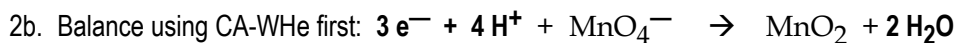
Then add enough OH^- to each side to neutralize H^+ .



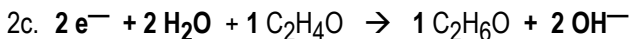
Then adjust the water.



Check: 8 O, 8 H, 1 Cr, 5 negative charges on each side. Balanced.



Check: 4 H, 6 O, 1 Mn, 4 negative charges on each side. Balanced.



* * * * *

Lesson 16B: Balancing By Adding Half-Reactions

Steps in Adding Half-Reactions

Two balanced *half*-reactions can be *added* to give a balanced *redox* equation.

The key principle in adding half-reactions is that the electrons lost must equal the electrons gained: the number of electrons in one half-reaction must equal the number of electrons on the *other* side of the arrow in the other half-reaction.

The steps to adding two half-reactions are

1. *Balance* each half-reaction. (The two balanced half-reactions that result must have their e^{-} terms on *opposite* sides; check your work if they do not.)
2. *Multiply* each half-reaction by a *number* (based on a lowest common multiple) to get the *same number of electrons* in both half-reactions.
3. *Add* the two half-reactions. All terms on the left side of the two arrows are added together on the left side of the arrow in the final, total reaction. All terms on the right are added on the right. Cancel "like terms" on each side. The like number of electrons on each side *must* cancel.
4. Check to make sure that the resulting *trial* redox equation is balanced for atoms and charge. If not, modify the final coefficients by trial and error.

Use those steps to balance and add these two half-reactions.

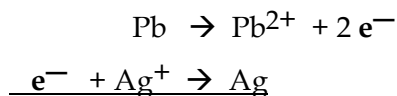


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Answer

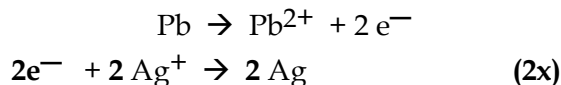
1. Balance each half-reaction separately for atoms and charge.

If possible, write the reactions so that the *arrows* line up, one below the other, to emphasize that we separately add the *two columns* separated by the arrows.

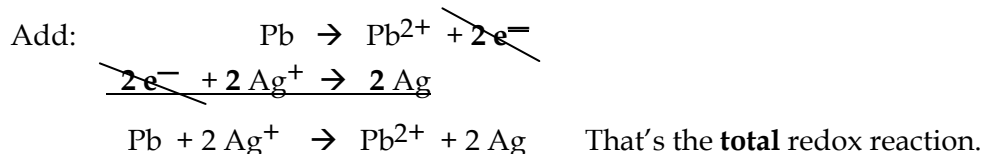


2. Multiply each half-reaction by a lowest common multiple (LCM) that results in the same number of electrons in both half-reactions.

Multiply all terms in the bottom equation by **2**. This keeps the half-reaction balanced, but makes the electron count equal but on opposite sides in the two equations..



3. Add the two half-reactions by adding terms that are on the same side of the arrows, then cancel like terms on each side. The electrons must cancel.



4. Check: 1 Pb, 2 Ag atoms on each side; +2 charge on each side. Balanced.

Summary: To add two half-reactions, make the *electron count the same* in both half-reactions (with the electrons on *opposite* sides).

* * * * *

Flashcards

One-way cards (with notch)

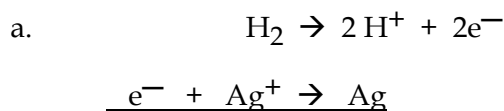
Back Side -- Answers

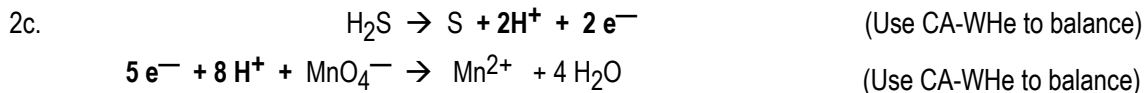
To balance a redox half-reaction, the steps are	CA-WHe!
In balancing a redox half-reaction, O atoms are balanced by adding	Water
In balancing a redox half-reaction, H atoms are balanced by adding	H ⁺
In balancing a redox half-reaction, charge is balanced by adding	e ⁻
To add two redox half-reactions the electrons must be	Equal in number and on opposite sides
To balance a half-reaction using OH ⁻ ions	Balance with CA-WHe, then neutralize H ⁺

Practice

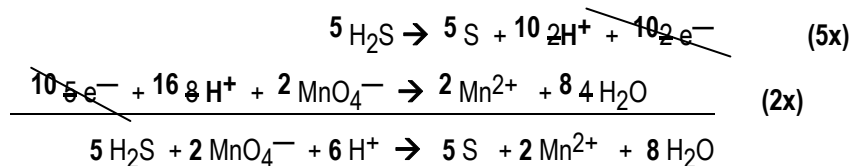
Try the last letter on each of these numbers. If easy, go to the next number. Need more practice? Do more.

1. Using the method above, add these half-reactions to get a balanced redox reaction.

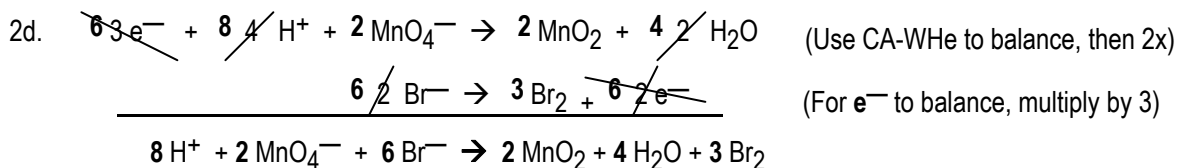




Use the LCM method to get electrons equal in both:

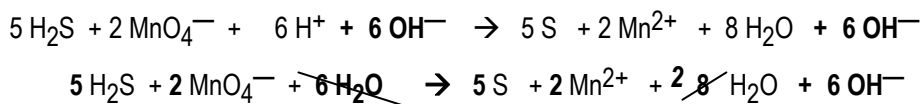


The 16 H⁺ on the **left** side of the arrow and the 10 H⁺ on the **right** cancel to give 6 H⁺ on the left. Check: 16 H, 5 S, 2 Mn, 8 O atoms on each side. Overall +4 charge on each side. Balanced.

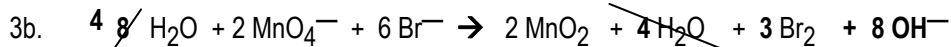


Check: 8 H, 2 Mn, 8 O, 6 Br atoms on each side. Overall zero charge on each side. Balanced.

3a. "In a basic solution" means using OH⁻ ions instead of H⁺. Neutralize the H⁺.



Check: 10 H, 5 S, 2 Mn, 8 O atoms on each side. Overall -2 charge on each side. Balanced.



Check: 8 H, 12 O, 2 Mn, 6 Br atoms on each side. Total -8 charge on each side. Balanced.

* * * * *

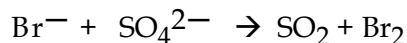
Lesson 16C: Separating Redox Into Half-Reactions

The fastest way to balance a redox reaction is often to separate the equation into two half-reactions. The steps are

1. Find the 2 atoms that *change* their oxidation numbers.
2. Write the 4 particles that contain those two atoms in *two* separate half-reactions, one above the other, with arrows lined up.
3. Balance and add the half-reactions.
4. Put the resulting trial coefficients back into the original equation. Adjust trial coefficients by trial and error if needed to balance atoms and charge.

Use those steps on the following problem.

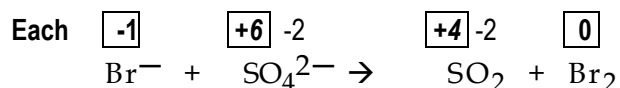
Q. Add two half-reactions, using H^+ and H_2O as needed, to balance this reaction.



Try Steps 1 and 2 above, and then check your answer below.

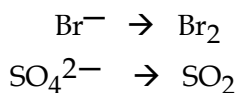
★ ★ ★ ★ ★

1. Find the two atoms that change their oxidation numbers.



Bromine and sulfur change their oxidation number.

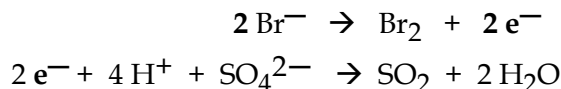
2. Write the 4 particles containing those two atoms in two separate half-reactions.



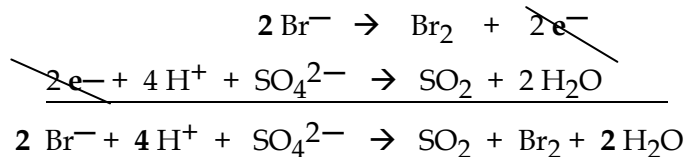
Now try Step 3 above, and then check your answer below.

★ ★ ★ ★ ★

3. Balance each half-reaction.



Since 2 electrons are in each, they can be added without multiplying either equation.



Try Step 4.

★ ★ ★ ★ ★

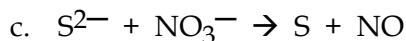
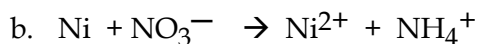
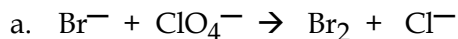
The unbalanced equation was $Br^- + SO_4^{2-} \rightarrow SO_2 + Br_2$

With H^+ , H_2O , and trial coefficients: $2 Br^- + 4 H^+ + SO_4^{2-} \rightarrow SO_2 + Br_2 + 2 H_2O$

Check the balancing. 4 H, 2 Br, 1 S, 4 O and zero net charge on each side.

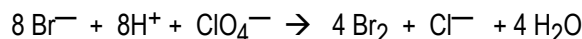
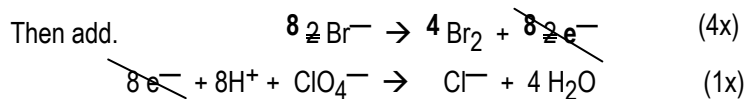
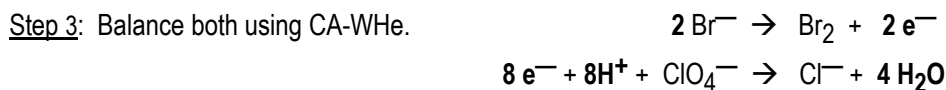
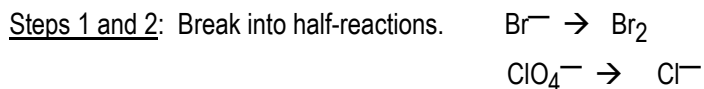
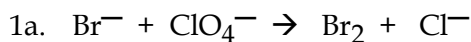
Practice

1. Balance these by adding half-reactions. Add H^+ ions and H_2O if needed.

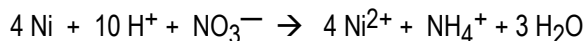
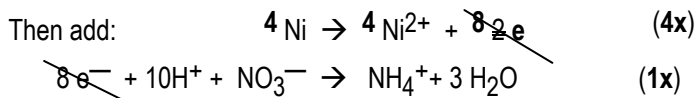
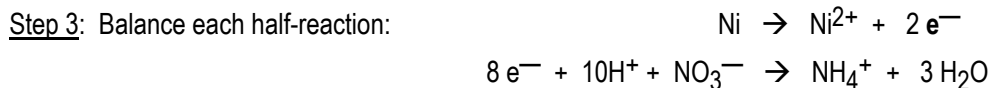
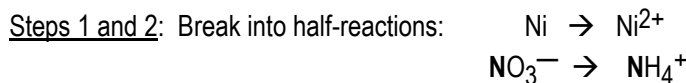
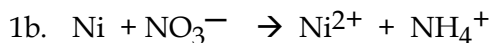


- In each reaction above, identify the reactant that is the reducing agent.
- In each reaction above, identify the reactant being reduced.
- Balance in basic solution: $\text{Cl}_2 \rightarrow \text{ClO}_3^- + \text{Cl}^-$

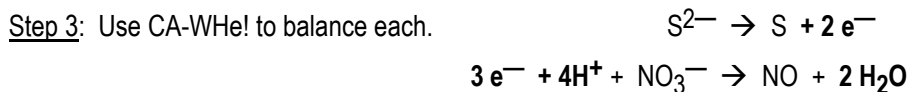
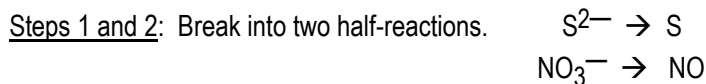
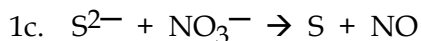
ANSWERS



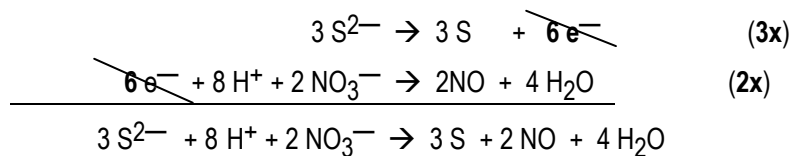
Step 4: Check: 8 Br, 8 H, 1 Cl, and 4 O atoms on each side; net -1 charge on each side.



Step 4: Check: 4 Ni, 10 H, 1 N, 3 O atoms on each side; $+9$ charge on each side.



Adjust the electrons to be the same on each side, then add the two half-reactions.



Step 4: 3 S, 8 H, 2 N, 6 O atoms on each side; zero net charge on each side. Check!

2. In problem 1a, the half-reactions show Br^- giving away electrons: acting as a reducing agent.

In problem 1b, the half-reactions show Ni giving away electrons: acting as a reducing agent. Metal atoms often act as reducing agents.

In problem 1c, the half reactions show the reactant S^{2-} donating electrons, which is the behavior of a reducing agent.

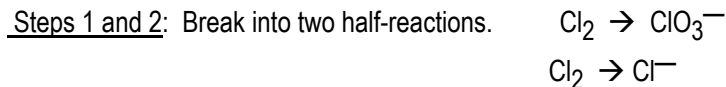
3. In problem 1a, the half-reactions show ClO_4^- gaining electrons, since the Cl atom is changing to a more negative oxidation number in the reaction. Gaining electrons means ClO_4^- is being reduced.

In problem 1b, the half-reactions show the N atom in the NO_3^- is having electrons added to it in the reaction: the NO_3^- is being reduced.

In problem 1c, the half-reactions show the N atom in the NO_3^- is again having electrons added to it in the reaction: the NO_3^- is being reduced.

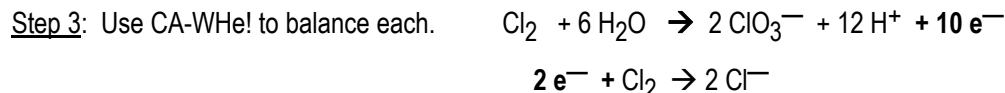
4. Balance in basic solution: $\text{Cl}_2 \rightarrow \text{ClO}_3^- + \text{Cl}^-$

"In a basic solution" means use OH^- ions. First balance using CA-WHe! method

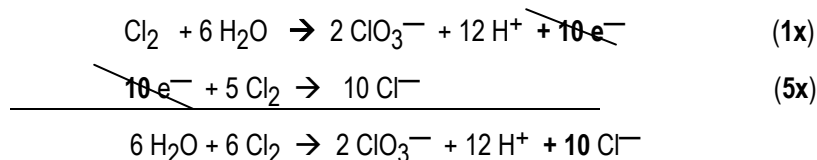


Cl_2 can both oxidize and reduce. If you needed that hint, adjust your work and finish.

* * * * *



Get the electrons to be the same on each side, and then add the two half-reactions.



Neutralize the H^+ : $12 \text{OH}^- + 6 \text{H}_2\text{O} + 6 \text{Cl}_2 \rightarrow 2 \text{ClO}_3^- + 12 \text{H}_2\text{O} + 10 \text{Cl}^-$

Adjust the H_2O : $12 \text{OH}^- + \cancel{6 \text{H}_2\text{O}} + 6 \text{Cl}_2 \rightarrow 2 \text{ClO}_3^- + \cancel{6} 12 \text{H}_2\text{O} + 10 \text{Cl}^-$

Reduce if possible: $6 \text{OH}^- + 3 \text{Cl}_2 \rightarrow 1 \text{ClO}_3^- + 3 \text{H}_2\text{O} + 5 \text{Cl}^-$

Step 4: 6 O, 6 H, 6 Cl atoms on each side; -6 charge net charge on each side. Check!

* * * * *

Lesson 16D: Balancing Redox With Spectator Ions Present

Balancing oxidation-reduction reactions can be tricky when spectator ions (ions that do not change in the reaction) are included in the equation. Consider this redox reaction carried out in an aqueous solution.



Balancing this equation by trial and error could take some time. Balancing using oxidation numbers can help, but for complex reactions, using half-reactions to balance is usually faster.

But when the spectators are present, how do you find the half-reactions? The following steps present a *system* to break complex redox reactions into half-reactions.

In your notebook, apply the following steps to the reaction above.

Steps for Balancing Redox With Spectators Present

1. The fundamental rule: To understand the reactions of ionic compounds, re-write the reaction using *separated-ion* formulas. You may leave out coefficients at this step.

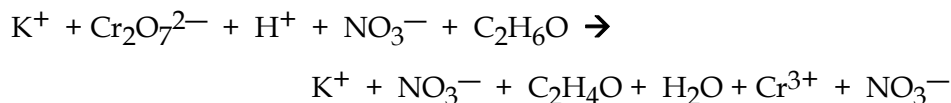
The rules for identifying *ionic* compounds (Lesson 7A), separating ionic *solid* into *separated-ion* formulas (Lesson 7C), and strong acid ionization (Lesson 14A) include:

- a. Compounds with both metal and non-metal atoms are usually ionic.
- b. Compounds containing alkali metal atoms (Li, Na, K, Rb, Cs, Fr) are soluble in water and separate ~100% to form monatomic +1 ions.
- c. Nitrates dissolve and ionize ~100% in water to form nitrate ions (NO_3^-).
- d. Aqueous solutions of strong acids (such as HCl and HNO_3) ionize ~100%.

Try Step 1 on the above reaction, then check your answer below.

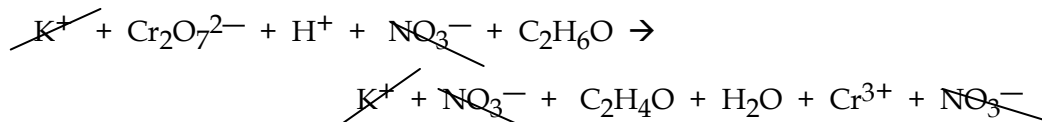
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The *separated-ions* version of the equation, leaving out coefficients, is



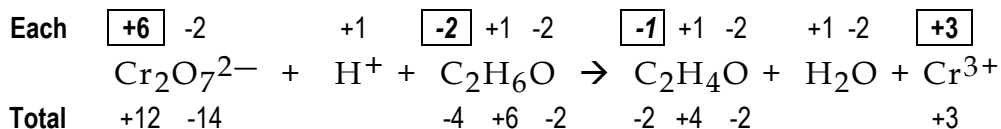
2. Cross out the spectators: particles that do not change in the reaction.

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3. With spectators omitted, find *two atoms* that *change oxidation number* in the reaction.

* * * * *



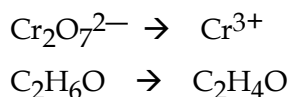
The two atoms that change oxidation number are Cr and C.

See if you can solve from here by splitting then adding the two half-reactions.

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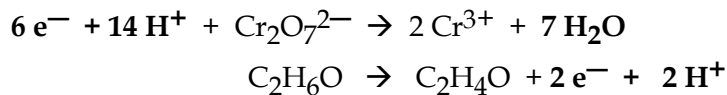
4. Split the 4 particles containing those two atoms into two half-reactions, writing one above the other.

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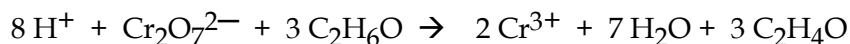
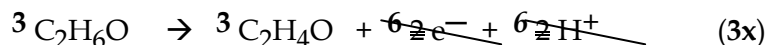
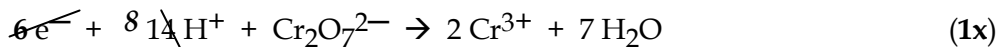
5. Balance the half-reactions using the CA-WHe! method.

* * * * *



6. Multiply the coefficients in each half-reaction to equalize the number of electrons in the two half-reactions, then add the half-reactions.

* * * * *



Check: 26 H, 2 Cr, 10 O, 6 C atoms on each side, +6 net charge on each side.

7. Plug the *trial* coefficients from the total of the half-reactions into the original equation. Then, using trial and error, finish balancing.

* * * * *



8. Check: 2K, 2 Cr, 34 O, 26 H, 8 N, 6 C on each side, neutral on each side.

In balancing, try to avoid calculator use.

If you find that you need a calculator to do the arithmetic of balancing and checking, make flashcards of the math facts on which you are rusty. Practicing those flashcards will give you the quick math recall that will help you to “keep your train of thought” during complex calculations.

Practice: First learn the steps above, then apply them to balance these redox reactions.

1. $\text{Cu} + \text{HNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{NO} + \text{H}_2\text{O}$
2. $\text{KMnO}_4 + \text{HNO}_3 + \text{H}_2\text{O}_2 \rightarrow \text{Mn}(\text{NO}_3)_2 + \text{O}_2 + \text{KNO}_3 + \text{H}_2\text{O}$
(Tip: H_2O_2 is hydrogen peroxide.)
3. $\text{MnO} + \text{PbO}_2 + \text{HCl} \rightarrow \text{HMnO}_4 + \text{PbCl}_2 + \text{H}_2\text{O}$ (Tip: HMnO_4 is a strong acid.)
4. In each reaction above, identify the reactant being oxidized.
5. In a basic solution: $\text{KClO}_3 + \text{KSH} \rightarrow \text{S} + \text{KCl}$

If you need additional practice, work the examples and/or problems that have supplied answers in any standard chemistry textbook.

Balancing Redox: Which Method To Use?

We have used three methods of balancing equations: the *trial and error* method (in Lesson 10B) that can be used on all equations, the *oxidation number* method for redox equations (in Lesson 15C), and the *half-reaction* method for redox in this module. Which method should you use?

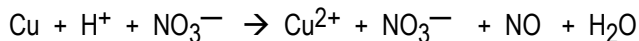
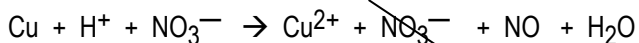
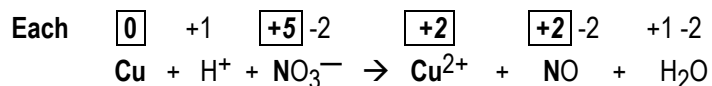
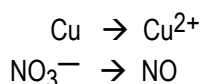
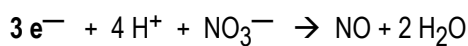
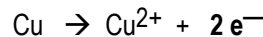
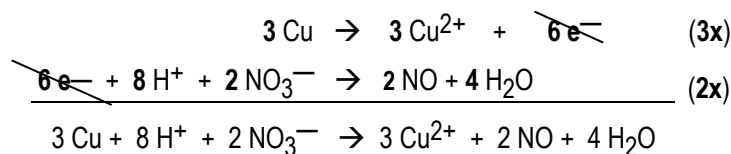
- *Trial and error* is always a legitimate method of balancing. However, for complex redox reactions, trial and error alone can be very time-consuming.

Two additional methods can be used to balance *redox* equations.

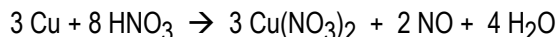
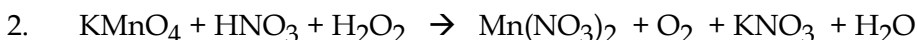
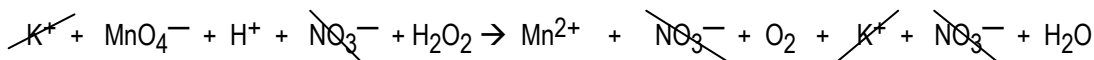
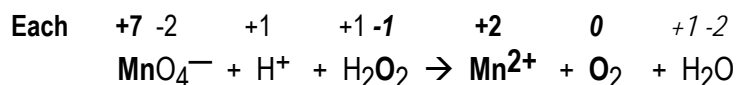
- The *oxidation-number* method supplies *four* key trial coefficients. The remaining coefficients must be determined by trial and error.
- The *half-reaction* method often takes longer at the start than the oxidation number method, but half-reactions nearly always supply *more* than 4 trial coefficients.

Which method is best? In general, for easy reactions, try trial and error first. For more complex redox, try oxidation numbers. For very complex redox, try half-reactions.

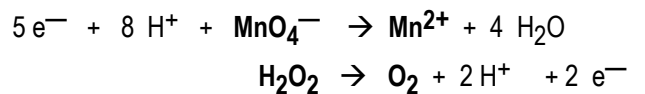
Remember, both redox methods give *trial* coefficients, not final answers. At the end of both methods, correct the coefficients if needed, using trial and error, until atoms and net charge are the same on each side.

ANSWERS**Step 1:** Break the compounds that ionize into *separated* ions. Leave out coefficients.**Step 2:** Cross out spectators ions: NO_3^- is a spectator on the right, but *not* on the left. Some NO_3^- on the left changed to NO, and some remained as NO_3^- .**Step 3:** Find the **two** atoms that change oxidation number.**Copper** and **nitrogen** are the central atoms that *change* their oxidation number.**Step 4:** Write the particles containing those two atoms in two separate half-reactions:**Step 5:** Complete half-reactions using CA-WHe!**Step 6:** Get the electrons equal in both, then add the two half-reactions. Use a lowest common multiple (LCM) method: the electron coefficient of one as the multiplier of the other:

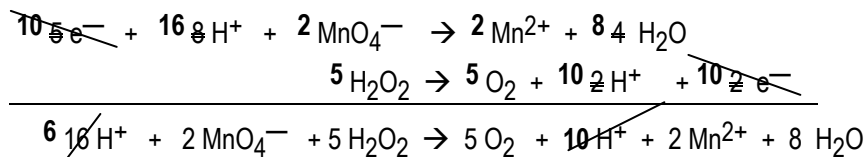
Check: 3 Cu, 8 H, 2 N, 6 O, +6 charge on each side.

Step 7: Plug the trial coefficients into the original equation; finish balancing by trial and error.**Step 8:** Check: 3 Cu, 8 H, 8 N, 24 O, zero net charge on each side. Balanced.**Steps 1 and 2:** Break compounds that ionize into *separated* ions, then cross out the spectators and particles that are the same on each side.**Step 3:** Find the 2 atoms that change their oxidation number. In *peroxides*, the O Ox# is -1.

Steps 4 and 5: Write the 4 particles containing the 2 atoms that change their oxidation number in separate half-reactions, then balance the half-reactions using the CA-WHe! method.

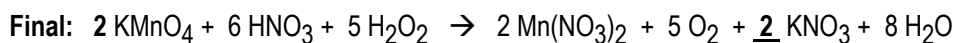
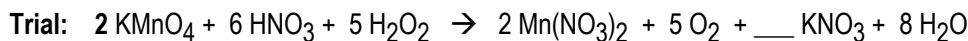


Step 6: Get electrons equal in both, then add the half-reactions.

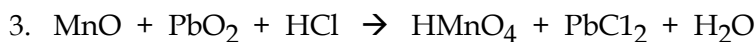


Check: 2 Mn, 18 O, 16 H atoms on each side, +4 charge on each side. Balanced.

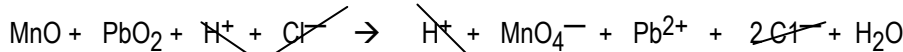
Step 7: Plug trial coefficients into the original equation; finish balancing by trial and error.



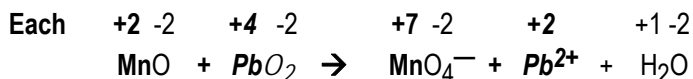
Step 8: 2 K, 2 Mn, 36 O, 16 H, 6 N atoms on each side, zero net charge on each side. Balanced.



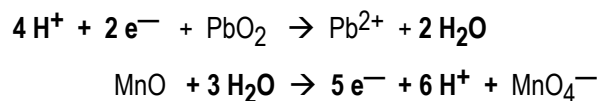
Steps 1 and 2: Break compounds that ionize into *separated* ions, then cross out the spectators and particles that are the same on each side: Strong acids ionize 100% in water.



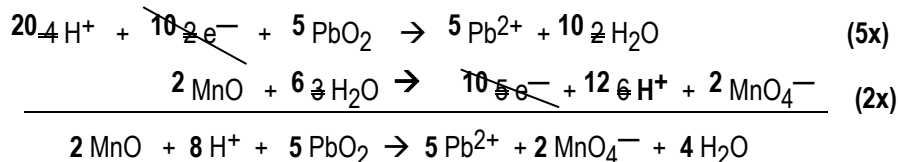
Step 3: Find the 2 atoms that change their oxidation number.



Steps 4 and 5: Write the 4 particles containing the 2 atoms that change their oxidation number in separate half-reactions, then balance the half-reactions using the CA-WHe! method.

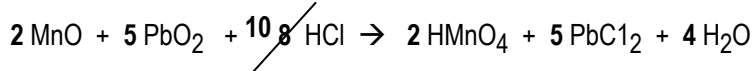


Step 6: Get electrons equal in both, then add the half-reactions.



Check: 2 Mn, 12 O, 8 H, 5 Pb atoms on each side, +8 charge on each side. Balanced.

Step 7: Plug trial coefficients into the original equation; finish balancing by trial and error.

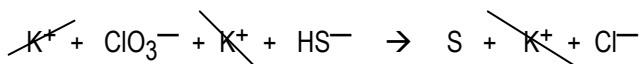


Step 8: Check: 2 Mn, 12 O, 10 H, 5 Pb, 10 Cl atoms, zero charge on each side. Balanced.

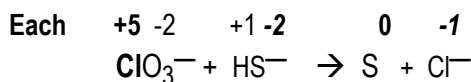
The HCl coefficient must be adjusted to balance the spectators, but the balanced half-reactions get you *close* to the final answer.

- In problem 1, the half-reactions show Cu losing electrons in the reaction. That means Cu is being oxidized.
In problem 2, the half reactions show the reactant H_2O_2 losing electrons -- being oxidized.
In problem 3, the half reactions show the reactant MnO losing electrons -- being oxidized.
- In a basic solution: $\text{KClO}_3 + \text{KSH} \rightarrow \text{S} + \text{KCl}$

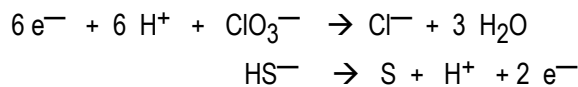
Steps 1 and 2: Break compounds that ionize into *separated* ions, then cross out particles that are the same on each side.



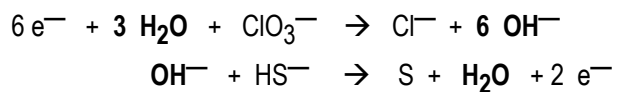
Step 3: Find the 2 atoms that change oxidation number.



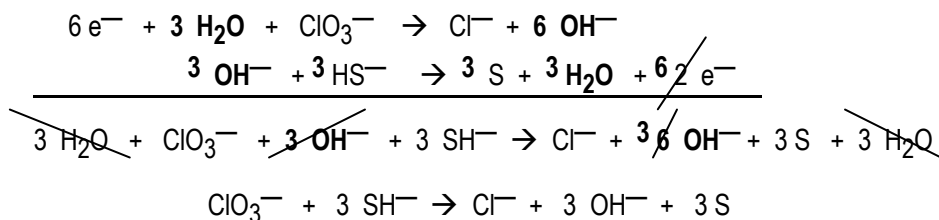
Steps 4 and 5: Write the 4 particles containing the 2 atoms that change oxidation number in separate half-reactions, then balance the half-reactions using the CA-WHE! method.



But "in a basic solution" means balance the half-reactions with OH^- ions. Neutralize the H^+ above.



Step 6: Get electrons equal in both, then add the half-reactions.



Check: 1 Cl, 3 O, 3 S, 3 H atoms on each side, -4 charge on each side. Balanced.

Step 7: Plug trial coefficients into the original equation; finish balancing by trial and error.



Step 8: 4 K, 1 Cl, 3 O, 3 S, 3 H atoms on each side, zero net charge on each side. Balanced.

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Lesson 16E: Review Quiz For Modules 13-16

You may use a calculator and a periodic table. Work on your own paper.

Set a 35-minute limit, then check your answers after the *Summary* that follows.

* * * * *

- (See Lesson 13A): Label each of these compounds as soluble or insoluble in water.
 - $(\text{NH}_4)_3\text{PO}_4$
 - AgBr
 - $\text{Pb}(\text{NO}_3)_2$
 - BaCO_3
 - CaCl_2
- (Lesson 13C): For the reaction $\text{Pb}(\text{NO}_3)_2 + \text{KCl} \rightarrow$
 - Write a total ionic equation.
 - Write the net ionic equation.
- (Lessons 12 E, 13D): In the reaction in Problem 2, if 0.100 L of $\text{Pb}(\text{NO}_3)_2$ solution is reacted with excess KCl , and the weight of the rinsed and dried solid product is 11.12 g, what was the original $[\text{Pb}(\text{NO}_3)_2]$?
- Supply the ratios (coefficients) for the two reactants: $\text{HNO}_3 + \text{Unk}(\text{OH})_3 \rightarrow$
- (Lesson 14D): If 912 mg of a dry solid acid is neutralized by 22.0 mL of 0.120 M NaOH , assuming that each acid particle contains two acidic hydrogens,
 - How many moles of acid were in the acid sample?
 - What is the molar mass of the unknown acid?
- (Lesson 14E): Write the final products in molecular (solid) formulas and balance this equation. Assume the reaction goes to completion.

$$\text{CH}_3\text{COOH} + \text{KHCO}_3 \rightarrow$$
- (Lesson 15D): Balance:

$$\text{FeCl}_2 + \text{KMnO}_4 + \text{HCl} \rightarrow \text{MnCl}_2 + \text{FeCl}_3 + \text{H}_2\text{O} + \text{KCl}$$
- (Lesson 15C): In problem 6, which substance is the reducing agent?
- (Lesson 15E): Based on Problem 6, how many grams of KMnO_4 (158.0 g/mol) are needed to react with 40.0 mL of 0.150 M iron(II) chloride?
- (Lesson 16A): Use water and protons to balance this half-reaction:

$$\text{IO}_3^- \rightarrow \text{I}^-$$
- (Lesson 16A): Balance this half-reaction in a basic solution.

$$\text{Br}_2 \rightarrow \text{BrO}^-$$

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Summary: Half-Reaction Balancing

1. Redox half-reactions can be constructed and balanced by using

The CA-WHe! Method

For redox reactions run in aqueous solutions, to balance half-reactions,

- a. First balance the *central atom* (CA), usually one that is *not* O or H. Then,
 - b. Add **W**ater if needed to balance **oxygen**.
 - c. Add **H**⁺ if needed to balance the **hydrogen**.
 - d. Add **e**lectrons to balance **charge**.
2. To balance half-reactions using OH⁻ instead H⁺ ions,
 - a. first balance by the CA-WHe method.
 - b. Then *neutralize* the H⁺ by adding OH⁻ ions equally to each side.
 - c. Adjust the H₂O coefficients on each side.
 3. Half-reactions can be added to balance redox reactions. The steps are
 - a. *Balance* each half-reaction.
 - b. *Multiply* each half-reaction by a lowest common multiple to get the *same number of electrons* in both half-reactions.
 - c. *Add* the two half-reactions. Cancel like terms on each side. A like number of electrons on each side **must** cancel.
 - d. *Check* to make sure that the resulting *trial* redox equation is balanced for atoms and charge.
 4. Redox reactions can be divided into half-reactions to aid in balancing.
 - a. Re-write the reaction changing any molecular (solid) into separated-ion formulas.
 - b. Find the two atoms that *change* their oxidation numbers in the reaction.
 - c. Write the 4 particles containing the atoms that change their oxidation numbers in *two* separate half-reactions.
 - d. Balance and add the half-reactions.
 - e. Check and adjust trial coefficients if needed to balance atoms and charge.

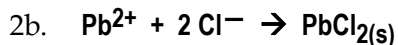
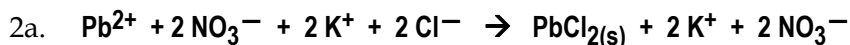
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ANSWERS - Module 13-16 Review Quiz

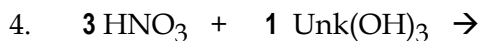
Some *partial* solutions are provided below. Your work on calculations should include WANTED, DATA, and SOLVE.

1a. **Soluble** 1b. **Insoluble**

1c. **Soluble** 1d. **Insoluble** 1e. **Soluble**

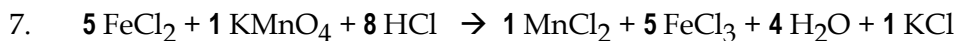
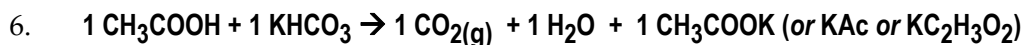


3. **0.400 M Pb(NO₃)₂** (0.04000 mol Pb(NO₃)₂ / 0.100 L Pb(NO₃) soln)



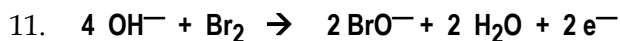
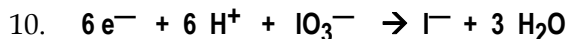
5a. **1.32 x 10⁻³ mol acid**

5b. **691 g/mol** (0.912 g acid / 1.32 x 10⁻³ mol acid)



8. **FeCl₂**

9. **0.190 g KMnO₄** ? g KMnO₄ = 0.0400 L FeCl₂ • $\frac{0.150 \text{ mol FeCl}_2}{1 \text{ L FeCl}_2 \text{ soln}}$ • $\frac{1 \text{ mol KMnO}_4}{5 \text{ mol FeCl}_2}$ • $\frac{158.0 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4}$ =



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