Redox: Half-Reaction Balancing: Self-Study Assignment

You will have a QUIZ on the attached pages on ____________________ .

Your assignment is: READ the pages attached. WORK the examples in the lesson. Complete the pages as homework.

To work the examples,

- use a sheet of paper to cover below the * * * * * line,
- try the problem on your paper,
- then check your answer below the * * * * * line.

Start early. This assignment will require 2-4 hours of work outside of class.

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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.

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

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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.

$$MnO_4^- + H^+ \rightarrow Mn^{2+} + H_2O$$

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

Both sides have 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:

$$MnO_4^- + \mathbf{8}\,H^+ \rightarrow Mn^{2+} + \mathbf{4}\,H_2O$$

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

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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.

$$\mathbf{5\ e^-} + MnO_4^- + 8\ H^+ \rightarrow Mn^{2+} + 4\ H_2O$$

The net charge is now **+2** on both sides. 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.

   a. $NO_3^- + 4\ H^+ \rightarrow NO + 2\ H_2O$

   b. $H_2O_2 \rightarrow 2\ H^+ + O_2$

2. Balance these half-reactions.

   a. SO_4^{2-} + H^+ $\rightarrow$ SO_2 + H_2O

   b. H_2S $\rightarrow$ S_8 + H^+

   c. MnO_2 + H^+ $\rightarrow$ Mn^{2+} + H_2O

   d. NH_4^+ + H_2O $\rightarrow$ NO_3^- + H^+

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 "**c**entral **a**tom" (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 both sides.

It helps to memorize: "To balance half-reactions, balance the *c*entral *a*tom, then WHe!"

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

Q. Balance this half-reaction.

$$Cr_2O_7^{2-} \rightarrow \quad Cr^{3+}$$

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Answer

1. First, balance the ***c**entral **a**tom* (usually an atom that is not O or H).

   $Cr_2O_7^{2-} \rightarrow \underline{\mathbf{2}}\ Cr^{3+}$

2. Add **W**ater to balance *oxygen* atoms.

   $Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + \mathbf{7\ H_2O}$

3. Add **H^+** to balance *hydrogen*.

   $\mathbf{14\ H^+} + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

4. Add **e**lectrons to balance *charge*.

   $\mathbf{6\ e^-} + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$

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

   Check: 14 H, 2 Cr, 7 O on both sides; **+6** charge on both sides. Balanced.

Practice B

Complete the balancing of these half-reactions, adding H^+ and H_2O 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.

1. $ClO_3^- \rightarrow Cl_2$
2. $PbSO_4 \rightarrow Pb + SO_4^{2-}$
3. $SO_2 \rightarrow SO_4^{2-}$
4. $MnO_4^- \rightarrow MnO_2$
5. $Br_2 \rightarrow BrO_4^-$
6. $Au \rightarrow Au^{3+}$
7. $H_2O \rightarrow O_2$

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 H_2O.
2. Then *neutralize* the H^+: using the H^+ coefficient, add that number of **OH^-** ions to *both* sides. This "neutralization of H^+" will replace the number of H^+ ions in the reaction

with an equal number of H_2O molecules, and add the same number of **OH^-** ions to the side opposite the original H^+.

3. Adjust the H_2O coefficients on both sides 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^+.

$$Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+}$$

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Answer

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

$$6\,e^- + 14\,H^+ + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$$

2. Neutralize the H^+ by adding OH^- to both sides.

$$6\,e^- + 14\,H^+ + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O \quad \mathbf{+\,14\,OH^-}$$
$$\mathbf{+\,14\,OH^-}$$

3. React the H^+ and OH^- to form H_2O.

$$6\,e^- + \mathbf{14\,H_2O} + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O \quad \mathbf{+\,14\,OH^-}$$

4. Adjust the count of the H_2O. $7\,H_2O$ cancel on both sides.

$$6\,e^- + \underset{\mathbf{7}}{\cancel{\mathbf{14}}}\,\mathbf{H_2O} + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + \cancel{7\,H_2O} \quad \mathbf{+\,14\,OH^-}$$

5. Check: 14 H, 2 Cr, 14 O, 8 negative charges on both sides.

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^+.

   a. $3\,e^- + NO_3^- + 4\,H^+ \rightarrow NO + 2\,H_2O$

   b. $H_2O_2 \rightarrow 2\,H^+ + O_2 + 2\,e^-$

2. Balance these half-reactions using OH^- and H_2O.

   a. $Cr^{3+} \rightarrow CrO_4^{2-}$

b. MnO_4^{-} → MnO_2

c. C_2H_4O → C_2H_6O

ANSWERS

Pretest

1. The gain of electrons. 2. $\mathbf{2e^{-}} + F_2 \rightarrow 2\,F^{-}$ 3. Cl = +5, Mn = +7

4a. Sn^{4+} (OA) + Co (RA) → 4b. Ca (RA) + 2 H^+ (OA) →

Practice A

The total of the charges on both sides must balance.

1a. $\mathbf{3\,e^{-}} + NO_3^{-} + 4\,H^+ \rightarrow NO + 2\,H_2O$ **1b**. $H_2O_2 \rightarrow 2\,H^+ + O_2 + \mathbf{2\,e^{-}}$

2a. $\mathbf{2\,e^{-}} + SO_4^{2-} + \mathbf{4}\,H^+ \rightarrow SO_2 + \mathbf{2}\,H_2O$ **2b.** $\mathbf{8}\,H_2S \rightarrow S_8 + \mathbf{16}\,H^+ + \mathbf{16\,e^{-}}$

2c. $MnO_2 + \mathbf{4}\,H^+ + \mathbf{2\,e^{-}} \rightarrow Mn^{2+} + \mathbf{2}\,H_2O$ **2d.** $\mathbf{1}\,NH_4^+ + \mathbf{3}\,H_2O \rightarrow \mathbf{1}\,NO_3^{-} + \mathbf{10}\,H^+ + \mathbf{8\,e^{-}}$

Practice B

1. First balance the Central Atom: $\mathbf{2}\,ClO_3^{-} \rightarrow Cl_2$

   Then WHe: Add <u>w</u>ater to balance **O**: $2\,ClO_3^{-} \rightarrow Cl_2 + \mathbf{6\,H_2O}$

   <u>$\mathbf{H^+}$</u> to balance **H:** $2\,ClO_3^{-} + \mathbf{12\,H^+} \rightarrow Cl_2 + 6\,H_2O$

   <u>$\mathbf{e^{-}}$</u> to balance charge**:** $2\,ClO_3^{-} + 12\,H^+ + \mathbf{10\,e^{-}} \rightarrow Cl_2 + 6\,H_2O$

   Check: 2 Cl, 6 O, 12 H, zero charge on both sides. Balanced.

2. Central atom, O, and H are already balanced. Use $\mathbf{e^{-}}$ to balance charge**.**

   $PbSO_4 + \mathbf{2\,e^{-}} \rightarrow Pb + SO_4^{2-}$

   Check: 1 Pb, 1 S, 4 O, 2 negative charges on both sides. Balanced.

3. Central atom S is balanced. Use **w**ater to balance O: $SO_2 + \mathbf{2\,H_2O} \rightarrow SO_4^{2-}$

   Use $\mathbf{H^+}$ to balance **H:** $SO_2 + 2\,H_2O \rightarrow 4H^+ + SO_4^{2-}$

   Use $\mathbf{e^{-}}$ to balance charge**:** $SO_2 + 2\,H_2O \rightarrow \mathbf{2\,e^{-}} + 4H^+ + SO_4^{2-}$

   Check: 1 S, 4 O, 4 H, zero charge on both sides. Balanced.

4. $MnO_4^{-} + \mathbf{4\,H^+} + \mathbf{3\,e^{-}} \rightarrow MnO_2 + \mathbf{2\,H_2O}$ 5. $Br_2 + \mathbf{8\,H_2O} \rightarrow \mathbf{2}\,BrO_4^{-} + \mathbf{16\,H^+} + \mathbf{14\,e^{-}}$

6. $Au \rightarrow Au^{3+} + \mathbf{3\,e^{-}}$ 7. $\mathbf{2}\,H_2O \rightarrow O_2 + \mathbf{4\,H^+} + \mathbf{4\,e^{-}}$

Practice C

1a. Add enough OH^- to both sides to neutralize H^+.

$$3\,e^- + NO_3^- + \cancel{4\,H^+} \rightarrow NO + 2\,H_2O \mathbf{\,+\,4\,OH^-}$$
$$\mathbf{4\,H_2O}$$

Adjust the H_2O $\quad 3\,e^- + NO_3^- + \mathbf{2\,\cancel{4}\,H_2O} \rightarrow NO + \cancel{2\,H_2O} + \mathbf{4\,OH^-}$

Final. $\mathbf{3\,e^- + NO_3^- + 2\,H_2O \rightarrow NO + 4\,OH^-}$

Check: 1 N, 5 O, 4H, 4 negative charges on both sides. Balanced.

1b. $\mathbf{2\,OH^-} + H_2O_2 \rightarrow \cancel{2\,H^+} + O_2 + 2\,e^-$
$$\mathbf{2\,H_2O}$$

Final: $\mathbf{2\,OH^- + H_2O_2 \rightarrow 2\,H_2O + O_2 + 2\,e^-}$

Check: 4 O, 4 H, 2 negative charges on both sides. Balanced.

2a. First balance using CA-WHe! method.

$$\mathbf{4\,H_2O} + Cr^{3+} \rightarrow CrO_4^{2-} \mathbf{\,+\,8\,H^+ + 3\,e^-}$$

Then add enough OH^- to both sides to neutralize H^+.

$$\mathbf{8\,OH^-} + 4\,H_2O + Cr^{3+} \rightarrow CrO_4^{2-} + \cancel{8\,H^+} + 3\,e^-$$
$$\mathbf{8\,H_2O}$$

Then adjust the water.

$$8\,OH^- + \cancel{\mathbf{4\,H_2O}} + Cr^{3+} \rightarrow CrO_4^{2-} + \mathbf{4\,\cancel{8}\,H_2O} + 3\,e^-$$

Final: $\mathbf{8\,OH^- + Cr^{3+} \rightarrow CrO_4^{2-} + 4\,H_2O + 3\,e^-}$

Check: 8 O, 8 H, 1 Cr, 5 negative charges on both sides. Balanced.

2b. Balance using CA-WHe first: $\mathbf{3\,e^- + 4\,H^+} + MnO_4^- \rightarrow MnO_2 + \mathbf{2\,H_2O}$

Neutralize the H^+: $3\,e^- + \cancel{\mathbf{4\,H^+}} \mathbf{\,+\,4\,H_2O} + MnO_4^- \rightarrow MnO_2 + 2\,H_2O \mathbf{\,+\,4\,OH^-}$

Adjust the H_2O: $\quad 3\,e^- \mathbf{\,+\,2\,\cancel{4}\,H_2O} + MnO_4^- \rightarrow MnO_2 + \cancel{\mathbf{2\,H_2O}} + 4\,OH^-$

Check: 4 H, 6 O, 1 Mn, 4 negative charges on both sides. Balanced.

2c. $\mathbf{2\,e^- + 2\,H_2O} + \mathbf{1}\,C_2H_4O \rightarrow \mathbf{1}\,C_2H_6O \mathbf{\,+\,2\,OH^-}$

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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 $\mathbf{e^-}$ terms on *opposite* sides; check your work if they do not.)
2. *Multiply* each half-reaction by a *number* (based on a lowest common denominator) 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 both sides. 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.

Q. Add:

$$Pb \rightarrow Pb^{2+}$$
$$\underline{Ag^+ \rightarrow Ag}$$

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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.

$$Pb \rightarrow Pb^{2+} + 2\,\mathbf{e^-}$$
$$\underline{\mathbf{e^-} + Ag^+ \rightarrow Ag}$$

2. Multiply each half-reaction by a lowest common denominator (LCD) to get 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..

$$Pb \rightarrow Pb^{2+} + 2\,e^-$$
$$\mathbf{2e^-} + \mathbf{2}\,Ag^+ \rightarrow \mathbf{2}\,Ag \qquad \mathbf{(2x)}$$

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

Add: $Pb \rightarrow Pb^{2+} + \cancel{\mathbf{2}\,e^-}$

$\cancel{\mathbf{2}\,e^-} + \mathbf{2}\,Ag^+ \rightarrow \mathbf{2}\,Ag$

$Pb + 2\,Ag^+ \rightarrow Pb^{2+} + 2\,Ag$ That's the **total** redox reaction.

4. Check: 1 Pb, 2 Ag atoms on both sides; +2 charge on both sides. Balanced.

Summary: To *add* two half-reactions, the *electron count* must be the *same* in both half-reactions, and the electrons must be on *opposite* sides.

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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	**W**ater
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.

a. $H_2 \rightarrow 2\,H^+ + 2e^-$

$e^- + Ag^+ \rightarrow Ag$

b. $2Cl^- \rightarrow Cl_2 + 2e^-$

$5\,e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4\,H_2O$

2. Balance each half-reaction, adding H^+ and H_2O if needed, then add the half-reactions to get a balanced redox reaction.

a. $Ni \rightarrow Ni^{2+}$

$Cl_2 \rightarrow Cl^-$

b. $Fe^{3+} \rightarrow Fe$

$NO \rightarrow NO_3^-$

c. $H_2S \rightarrow S$

$MnO_4^- \rightarrow Mn^{2+}$

d. $MnO_4^- \rightarrow MnO_2$

$Br^- \rightarrow Br_2$

3. Balance the result in Problems 2c and 2d using OH^- ions instead of H^+.

ANSWERS

1a. $H_2 \rightarrow 2\ H^+ + 2e^-$

$2\ e^- + 2\ Ag^+ \rightarrow 2\ Ag$ (multiply all terms in this half-reaction by **2** so **e**⁻ balance)

$H_2 + 2\ Ag^+ \rightarrow 2\ H^+ + 2\ Ag$ (add terms on same side of arrows; check for balance)

1b. $10\ \cancel{2}Cl^- \rightarrow 5\ Cl_2 + \cancel{10\ \cancel{2}e^-}$ (To get **e**⁻ to balance, multiply by 5)

$\cancel{10\ \cancel{5}\ e^-} + 16\ \cancel{8}H^+ + 2\ MnO_4^- \rightarrow 2\ Mn^{2+} + 8\ \cancel{4}H_2O$ (and this one by 2)

$10\ Cl^- + 2\ MnO_4^- + 16\ H^+ \rightarrow 5\ Cl_2 + 2\ Mn^{2+} + 8\ H_2O$

Check: 10 Cl, 2 Mn, 8 O, 16 H atoms on both sides. **+4** charge on both sides. Balanced.

2a. $Ni \rightarrow Ni^{2+} + \cancel{2\ e^-}$ (Using CA-WHe, only need **e**⁻ to balance charge)

$\cancel{2\ e^-} + Cl_2 \rightarrow 2\ Cl^-$ (Balance central atom, then add **e**⁻ to balance charge)

$\mathbf{Ni + Cl_2 \rightarrow Ni^{2+} + 2\ Cl^-}$ (Since **2 e**⁻ on both sides, half-reactions can be added)

Check balance of atoms and charge: 1 Ni, 2 Cl atoms on both sides, neutral on both sides.

2b. $Fe^{3+} + \cancel{3\ e^-} \rightarrow Fe$ (Using CA-WHe, need only electrons to balance charge)

$NO + 2\ H_2O \rightarrow NO_3^- + 4\ H^+ + \cancel{3\ e^-}$ (Use CA-WHe to balance)

$\mathbf{Fe^{3+} + NO + 2H_2O \rightarrow Fe + NO_3^- + 4H^+}$ (3 **e**⁻ on both sides, halves can be added)

Check: 1 Fe, 1N, 3 O, 4 H atoms on both sides; +3 charge on both sides. Balanced.

2c. $H_2S \rightarrow S + 2H^+ + 2\ e^-$ (Use CA-WHe to balance)

$5\ e^- + 8\ H^+ + MnO_4^- \rightarrow Mn^{2+} + 4\ H_2O$ (Use CA-WHe to balance)

Use LCD method to get electrons equal in both:

$5\ H_2S \rightarrow 5\ S + \cancel{10\ \cancel{2}H^+ + 10\ \cancel{2}\ e^-}$ **(5x)**

$\cancel{10\ \cancel{5}\ e^-} + 16\ \cancel{8}\ H^+ + 2\ MnO_4^- \rightarrow 2\ Mn^{2+} + 8\ \cancel{4}\ H_2O$ **(2x)**

$5\ H_2S + 2\ MnO_4^- + 6\ H^+ \rightarrow 5\ S + 2\ Mn^{2+} + 8\ H_2O$

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 both sides. Overall +4 charge on both sides. Balanced.

2d. $\mathbf{6}\,\cancel{3\,e^-} + \mathbf{8}\,\cancel{4}\,H^+ + \mathbf{2}\,MnO_4^- \rightarrow \mathbf{2}\,MnO_2 + \mathbf{4}\,\cancel{2}\,H_2O$ (Use CA-WHe to balance, then 2x)

$\mathbf{6}\,\cancel{2}\,Br^- \rightarrow \mathbf{3}\,Br_2 + \cancel{\mathbf{6}\,2\,e^-}$ (For **e⁻** to balance, multiply by 3)

$\mathbf{8}\,H^+ + \mathbf{2}\,MnO_4^- + \mathbf{6}\,Br^- \rightarrow \mathbf{2}\,MnO_2 + \mathbf{4}\,H_2O + \mathbf{3}\,Br_2$

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

3a. $5\,H_2S + 2\,MnO_4^- + 6\,H^+ \mathbf{+ 6\,OH^-} \rightarrow 5\,S + 2\,Mn^{2+} + 8\,H_2O \mathbf{+ 6\,OH^-}$

$\mathbf{5}\,H_2S + \mathbf{2}\,MnO_4^- + \cancel{\mathbf{6\,H_2O}} \rightarrow \mathbf{5}\,S + \mathbf{2}\,Mn^{2+} + \mathbf{2}\,\cancel{8}\,H_2O \mathbf{+ 6\,OH^-}$

Check: 10 H, 5 S, 2 Mn, 8 O atoms on both sides. Overall —2 charge on both sides. Balanced.

3b. $\mathbf{4}\,\cancel{8}\,H_2O + 2\,MnO_4^- + 6\,Br^- \rightarrow 2\,MnO_2 + \cancel{\mathbf{4}\,H_2O} + \mathbf{3}\,Br_2 \mathbf{+ 8\,OH^-}$

Check: 8 H, 12 O, 2 Mn, 6 Br atoms on both sides. Total — 8 charge on both sides. Balanced.

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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, adding H^+ and H_2O if needed, to balance this reaction.

$$Br^- + SO_4^{2-} \rightarrow SO_2 + Br_2$$

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

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

$$\begin{array}{lccccc} \textbf{Each} & \boxed{-1} & & \boxed{+6}\ -2 & \boxed{+4}\ -2 & \boxed{0} \\ & Br^- & + & SO_4^{2-} \rightarrow & SO_2\ + & Br_2 \end{array}$$

Bromine and sulfur change their oxidation number.

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

$$Br^- \rightarrow Br_2$$
$$SO_4^{2-} \rightarrow SO_2$$

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

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3. Balance each half-reaction.

$$\mathbf{2}\,Br^{-} \rightarrow Br_2 + \mathbf{2\,e^{-}}$$

$$2\,\mathbf{e}^{-} + 4\,H^{+} + SO_4^{2-} \rightarrow SO_2 + 2\,H_2O$$

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

$$\mathbf{2}\,Br^{-} \rightarrow Br_2 + \cancel{2\,\mathbf{e}^{-}}$$

$$\cancel{2\,\mathbf{e}^{-}} + 4\,H^{+} + SO_4^{2-} \rightarrow SO_2 + 2\,H_2O$$

$$\mathbf{2}\,Br^{-} + \mathbf{4}\,H^{+} + SO_4^{2-} \rightarrow SO_2 + Br_2 + \mathbf{2}\,H_2O$$

Try Step 4.

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

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

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.

   a. $Br^{-} + ClO_4^{-} \rightarrow Br_2 + Cl^{-}$

   b. $Ni + NO_3^{-} \rightarrow Ni^{2+} + NH_4^{+}$

   c. $S^{2-} + NO_3^{-} \rightarrow S + NO$

2. In each reaction above, identify the reactant that is the reducing agent.
3. In each reaction above, identify the reactant being reduced.

ANSWERS

1a. $Br^{-} + ClO_4^{-} \rightarrow Br_2 + Cl^{-}$

<u>Steps 1 and 2</u>: Break into half-reactions. $Br^{-} \rightarrow Br_2$

$ClO_4^{-} \rightarrow Cl^{-}$

<u>Step 3</u>: Balance both using CA-WHe. $\mathbf{2}\,Br^{-} \rightarrow Br_2 + \mathbf{2\,e^{-}}$

$\mathbf{8\,e^{-}} + \mathbf{8H^{+}} + ClO_4^{-} \rightarrow Cl^{-} + \mathbf{4\,H_2O}$

Then add.

$$\mathbf{8}\,\cancel{2}\,Br^{-} \rightarrow \mathbf{4}\,Br_2 + \cancel{\mathbf{8}\,\cancel{2}\,\mathbf{e}^{-}} \quad (4x)$$

$$\cancel{8\,e^{-}} + 8H^{+} + ClO_4^{-} \rightarrow Cl^{-} + 4\,H_2O \quad (1x)$$

$$8\,Br^{-} + 8H^{+} + ClO_4^{-} \rightarrow 4\,Br_2 + Cl^{-} + 4\,H_2O$$

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

1b. $Ni + NO_3^- \rightarrow Ni^{2+} + NH_4^+$

Steps 1 and 2: Break into half-reactions:

$Ni \rightarrow Ni^{2+}$

$\mathbf{N}O_3^- \rightarrow \mathbf{N}H_4^+$

Step 3: Balance each half-reaction:

$Ni \rightarrow Ni^{2+} + 2\,\mathbf{e^-}$

$8\,e^- + 10H^+ + NO_3^- \rightarrow NH_4^+ + 3\,H_2O$

Then add:

$\mathbf{4}\,Ni \rightarrow \mathbf{4}\,Ni^{2+} + \cancel{\mathbf{8}\ 2\,\mathbf{e}}$ (**4x**)

$\cancel{8\,e^-} + 10H^+ + NO_3^- \rightarrow NH_4^+ + 3\,H_2O$ (**1x**)

$4\,Ni + 10\,H^+ + NO_3^- \rightarrow 4\,Ni^{2+} + NH_4^+ + 3\,H_2O$

Step 4: Check: 4 Ni, 10 H, 1 N, 3 O atoms on both sides; **+9** charge on both sides.

1c. $S^{2-} + NO_3^- \rightarrow S + NO$

Steps 1 and 2: Break into two half-reactions.

$S^{2-} \rightarrow S$

$NO_3^- \rightarrow NO$

Step 3: Use CA-WHe! to balance each.

$S^{2-} \rightarrow S\ \mathbf{+\,2\,e^-}$

$\mathbf{3\,e^-\ +\,4H^+} + NO_3^- \rightarrow NO + \mathbf{2\,H_2O}$

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

$3\,S^{2-} \rightarrow 3\,S + \cancel{\mathbf{6\,e^-}}$ (**3x**)

$\cancel{\mathbf{6}\,e^-} + 8\,H^+ + 2\,NO_3^- \rightarrow 2NO + 4\,H_2O$ (**2x**)

$3\,S^{2-} + 8\,H^+ + 2\,NO_3^- \rightarrow 3\,S + 2\,NO + 4\,H_2O$

Step 4: 3 S, 8H, 2 N, 6 O atoms on both sides; zero net charge on both sides. 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.

* * * * *

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.

$$K_2Cr_2O_7 + HNO_3 + C_2H_6O \rightarrow KNO_3 + C_2H_4O + H_2O + Cr(NO_3)_3$$

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.

* * * * *

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

$$K^+ + Cr_2O_7^{2-} + H^+ + NO_3^- + C_2H_6O \rightarrow$$
$$K^+ + NO_3^- + C_2H_4O + H_2O + Cr^{3+} + NO_3^-$$

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

* * * * *

$$\cancel{K^+} + Cr_2O_7^{2-} + H^+ + \cancel{NO_3^-} + C_2H_6O \rightarrow$$
$$\cancel{K^+} + \cancel{NO_3^-} + C_2H_4O + H_2O + Cr^{3+} + \cancel{NO_3^-}$$

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

* * * * *

	$Cr_2O_7^{2-}$	+	H^+	+	C_2H_6O	$\rightarrow$	C_2H_4O	+	H_2O	+	Cr^{3+}
Each	$\boxed{+6}$ -2		+1		$\boxed{-2}$ +1 -2		$\boxed{-1}$ +1 -2		+1 -2		$\boxed{+3}$
Total	+12 -14				-4 +6 -2		-2 +4 -2				+3

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.

* * * * *

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

* * * * *

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$
$$C_2H_6O \rightarrow C_2H_4O$$

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

* * * * *

$$\mathbf{6\ e^- + 14\ H^+} + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + \mathbf{7\ H_2O}$$
$$C_2H_6O \rightarrow C_2H_4O + \mathbf{2\ e^- + 2\ H^+}$$

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

* * * * *

$$\cancel{6\ e^-} + {}^{8}\cancel{14}\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O \qquad (\mathbf{1x})$$
$${}^{3}\ C_2H_6O \rightarrow {}^{3}\ C_2H_4O + \cancel{{}^{6}\cancel{2}\ e^- + {}^{6}\cancel{2}\ H^+} \qquad (\mathbf{3x})$$

$$8\ H^+ + Cr_2O_7^{2-} + 3\ C_2H_6O \rightarrow 2\ Cr^{3+} + 7\ H_2O + 3\ C_2H_4O$$

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

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

* * * * *

$$1\ K_2Cr_2O_7 + 8\ HNO_3 + 3\ C_2H_6O \rightarrow 2\ KNO_3 + 3\ C_2H_4O + 7\ H_2O + 2\ Cr(NO_3)_3$$

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

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" in upcoming more complex calculations.

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

1. $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$
2. $KMnO_4 + HNO_3 + H_2O_2 \rightarrow Mn(NO_3)_2 + O_2 + KNO_3 + H_2O$
   (Tip: H_2O_2 is hydrogen peroxide.)
3. $MnO + PbO_2 + HCl \rightarrow HMnO_4 + PbCl_2 + H_2O$ (Tip: $HMnO_4$ is an strong acid.)
4. In each reaction above, identify the reactant being oxidized.

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 both sides.

ANSWERS

1. $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$

   Step 1: Break the compounds that ionize into *separated* ions. Leave out coefficients.

   $Cu + H^+ + NO_3^- \rightarrow Cu^{2+} + NO_3^- + NO + H_2O$

   Step 2: Cross out spectators ions:

$$Cu + H^+ + NO_3^- \rightarrow Cu^{2+} + \cancel{NO_3^-} + NO + H_2O$$

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.

$$\textbf{Each} \quad \boxed{0} \quad +1 \quad \boxed{+5}\ -2 \qquad \boxed{+2} \qquad \boxed{+2}\ -2 \qquad +1\ -2$$

$$\mathbf{Cu} + H^+ + \mathbf{N}O_3^- \rightarrow \mathbf{Cu}^{2+} + \mathbf{N}O + H_2O$$

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:

$$Cu \rightarrow Cu^{2+}$$
$$NO_3^- \rightarrow NO$$

Step 5: Complete half-reactions using CA-WHe!

$$Cu \rightarrow Cu^{2+} + \mathbf{2\ e^-}$$
$$\mathbf{3\ e^-} + 4\ H^+ + NO_3^- \rightarrow NO + 2\ H_2O$$

Step 6: Get the electrons equal in both, then add the two half-reactions. Use a lowest common denominator (LCD) method: the electron coefficient of one as the multiplier of the other:

$$\mathbf{3}\ Cu \rightarrow \mathbf{3}\ Cu^{2+} + \cancel{\mathbf{6\ e^-}} \qquad (\mathbf{3x})$$
$$\cancel{\mathbf{6\ e^-}} + \mathbf{8}\ H^+ + \mathbf{2}\ NO_3^- \rightarrow \mathbf{2}\ NO + \mathbf{4}\ H_2O \qquad (\mathbf{2x})$$

$$3\ Cu + 8\ H^+ + 2\ NO_3^- \rightarrow 3\ Cu^{2+} + 2\ NO + 4\ H_2O$$

Check: 3 Cu, 8 H, 2 N, 6 O, +6 charge on both sides.

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

$$3\ Cu + 8\ HNO_3 \rightarrow 3\ Cu(NO_3)_2 + 2\ NO + 4\ H_2O$$

Step 8: Check: 3 Cu, 8 H, 8 N, 24 O, zero net charge on both sides. Balanced.

2. $KMnO_4 + HNO_3 + H_2O_2 \rightarrow Mn(NO_3)_2 + O_2 + KNO_3 + H_2O$

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

$$\cancel{K^+} + MnO_4^- + H^+ + \cancel{NO_3^-} + H_2O_2 \rightarrow Mn^{2+} + \cancel{NO_3^-} + O_2 + \cancel{K^+} + \cancel{NO_3^-} + H_2O$$

Step 3: Find the 2 atoms that change their oxidation number. In *per*oxides, the O Ox# is —1.

$$\textbf{Each} \quad \mathbf{+7}\ -2 \qquad +1 \qquad +1\ \boldsymbol{-1} \qquad \mathbf{+2} \qquad \boldsymbol{0} \qquad +1\ -2$$

$$\mathbf{Mn}O_4^- + H^+ + H_2\mathbf{O_2} \rightarrow \mathbf{Mn^{2+}} + \mathbf{O_2} + H_2O$$

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.

$$5\ e^- + 8\ H^+ + \mathbf{MnO_4^-} \rightarrow \mathbf{Mn^{2+}} + 4\ H_2O$$
$$\mathbf{H_2O_2} \rightarrow \mathbf{O_2} + 2\ H^+ + 2\ e^-$$

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

$$\cancel{10\ \cancel{5}\ e^-} + 16\ \cancel{8}\ H^+ + 2\ MnO_4^- \rightarrow 2\ Mn^{2+} + 8\ \cancel{4}\ H_2O$$

$$5\ H_2O_2 \rightarrow 5\ O_2 + 10\ \cancel{2}\ H^+ + \cancel{10\ \cancel{2}\ e^-}$$

$$6\ \cancel{16}\ H^+ + 2\ MnO_4^- + 5\ H_2O_2 \rightarrow 5\ O_2 + \cancel{10\ H^+} + 2\ Mn^{2+} + 8\ H_2O$$

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

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

Trial: $\mathbf{2}\ KMnO_4 + 6\ HNO_3 + 5\ H_2O_2 \rightarrow 2\ Mn(NO_3)_2 + 5\ O_2 + \_\_\_\ KNO_3 + 8\ H_2O$

Final: $\mathbf{2}\ KMnO_4 + 6\ HNO_3 + 5\ H_2O_2 \rightarrow 2\ Mn(NO_3)_2 + 5\ O_2 + \underline{\mathbf{2}}\ KNO_3 + 8\ H_2O$

<u>Step 8:</u> 2 K, 2 Mn, 36 O, 16 H, 6 N atoms on both sides, zero net charge on both sides. Balanced.

3. $MnO + PbO_2 + HCl \rightarrow HMnO_4 + PbCl_2 + H_2O$

<u>Steps 1 and 2</u>: Break compounds that ionize into *separated* ions, then cross out the spectators and particles that are the same on both sides: Strong acids ionize 100% in water.

$$MnO + PbO_2 + \cancel{H^+} + \cancel{Cl^-} \rightarrow \cancel{H^+} + MnO_4^- + Pb^{2+} + \cancel{2\ Cl^-} + H_2O$$

<u>Step 3</u>: Find the 2 atoms that change their oxidation number.

Each	**+2** -2		**+4** -2		**+7** -2		**+2**		+1 -2
	MnO	**+**	***PbO*$_2$**	$\rightarrow$	**MnO$_4^-$**	**+**	***Pb*$^{2+}$**	+	H_2O

<u>Steps 4 and 5:</u> 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.

$$\mathbf{4\ H^+ + 2\ e^-} + PbO_2 \rightarrow Pb^{2+} + \mathbf{2\ H_2O}$$

$$MnO + \mathbf{3\ H_2O} \rightarrow \mathbf{5\ e^-} + \mathbf{6\ H^+} + MnO_4^-$$

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

$$20\ \cancel{4}\ H^+ + \cancel{10\ \cancel{2}\ e^-} + 5\ PbO_2 \rightarrow 5\ Pb^{2+} + 10\ \cancel{2}\ H_2O \qquad \mathbf{(5x)}$$

$$2\ MnO + 6\ \cancel{3}\ H_2O \rightarrow \cancel{10\ \cancel{5}\ e^-} + 12\ \cancel{6}\ \mathbf{H^+} + 2\ MnO_4^- \qquad \mathbf{(2x)}$$

$$\mathbf{2}\ MnO + \mathbf{8}\ H^+ + \mathbf{5}\ PbO_2 \rightarrow \mathbf{5}\ Pb^{2+} + \mathbf{2}\ MnO_4^- + \mathbf{4}\ H_2O$$

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

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

$$\mathbf{2}\ MnO + \mathbf{5}\ PbO_2 + \mathbf{10}\ \cancel{8}\ HCl \rightarrow \mathbf{2}\ HMnO_4 + \mathbf{5}\ PbCl_2 + \mathbf{4}\ H_2O$$

<u>Step 8</u>: Check: 2 Mn, 12 O, 10 H, 5 Pb, 10 Cl atoms, zero charge on both sides. Balanced.

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

4. 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.

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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.

* * * * *

1. Label each of these compounds as soluble or insoluble in water.

   a. $(NH_4)_3PO_4$ b. AgBr c. $Pb(NO_3)_2$ d. $BaCO_3$ e. $CaCl_2$

2. For the reaction $Pb(NO_3)_2 + KCl \rightarrow$

   a. Write a total ionic equation. b. Write the net ionic equation.

3. In the reaction in Problem 2, if 0.100 L of $Pb(NO_3)_2$ solution is reacted with excess KCl, and the weight of the rinsed and dried solid product is 11.12 g, what was the original $[Pb(NO_3)_2]$?

4. 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,

   a. How many moles of acid were in the acid sample?

   b. What is the molar mass of the unknown acid?

5. Write the final products in molecular (solid) formulas and balance this equation. Assume the reaction goes to completion.

   $CH_3COOH + KHCO_3 \rightarrow$

6. Balance:

   $FeCl_2 + KMnO_4 + HCl \rightarrow MnCl_2 + FeCl_3 + H_2O + KCl$

7. In problem 6, which substance is the reducing agent?

8. 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?

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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 both sides.
   c. Adjust the H_2O coefficients on both sides.
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 denominator to get the *same number of electrons* in both half-reactions.
   c. *Add* the two half-reactions. Cancel like terms on both sides. 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.

* * * * *

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** (Lesson 13A)

2a. $\mathbf{Pb^{2+} + 2\ NO_3^- + 2\ K^+ + 2\ Cl^- \rightarrow PbCl_{2(s)} + 2\ K^+ + 2\ NO_3^-}$

2a. $\mathbf{Pb^{2+} + 2\ Cl^- \rightarrow PbCl_{2(s)}}$

3. **0.400 M $Pb(NO_3)_2$** (0.04000 mol $Pb(NO_3)_2$ / 0.100 L $Pb(NO_3)$ soln -- Lesson 13D)

4a. **1.32×10^{-3} mol acid** 4b. **691 g/mol** (0.912 g acid / 1.32×10^{-3} mol acid -- Lesson 14D)

5. **1 CH_3COOH + 1 $KHCO_3$ → 1 $CO_{2(g)}$ + 1 H_2O + 1 CH_3COOK (*or* KAc *or* $KC_2H_3O_2$)** (Lesson 14E)

6. **5** $FeCl_2$ + **1** $KMnO_4$ + **8** HCl → **1** $MnCl_2$ + **5** $FeCl_3$ + **4** H_2O + **1** KCl (Mods 15 and 16)

7. **$FeCl_2$** (Lesson 15B)

8. **0.190 g $KMnO_4$** (Lesson 15E)
$$? \text{ g } KMnO_4 = 0.0400 \text{ L } FeCl_2 \cdot \frac{0.150 \text{ mol } FeCl_2}{1 \text{ L } FeCl_2 \text{ soln}} \cdot \frac{1 \text{ mol } KMnO_4}{5 \text{ mol } FeCl_2} \cdot \frac{158.0 \text{ g } KMnO_4}{1 \text{ mol } KMnO_4} =$$

\# \# \# \# \#