Chem 170

Stoichiometric Calculations

Module Four

Balancing Chemical Reactions

Introduction to Module Four

When making a cheeseburger you might use a hamburger patty, cheese, an English muffin, pickles, onions, and mustard. We can represent this recipe symbolically as

hamburger patty + cheese + English muffin +
pickles + onions + mustard → cheeseburger

where the plus sign (+) means "combines with" and the arrow (→) means "yields" or "results in." Those items to the left of the arrow are the ingredients and the item to the right of the arrow is the final product.

Something important is missing, however, in this symbolic recipe for preparing a cheeseburger. When you make a cheeseburger, you want it to taste good. Specifically, you the cheeseburger to have the right amount of pickles, onions, and mustard to make it tasty. Adding coefficients before each ingredient

1 hamburger patty + 1 slice of cheese + 1 English muffin +
3 pickles + 2 slices of onion + 1 squirt of mustard → 1 yummy cheeseburger

gives a more complete symbolic recipe a cheeseburger. We call this symbolic recipe balanced because it specifies exactly how the ingredients are combined to make a cheeseburger.

In the same manner, we write balanced symbolic equations for chemical reactions. For example, propane, C_3H_8, burns in the presence of oxygen, forming carbon dioxide and water. We represent this reaction symbolically as

$$\underline{1}C_3H_8 + \underline{5}O_2 \rightarrow \underline{3}CO_2 + \underline{4}\ H_2O$$

where the plus sign means "reacts with" and the underlined numbers are the reaction's stoichiometric coefficients.[†] Species to the left of the arrow are called reactants and those to the right of the arrow are products. In this module, you will learn how to balance many types of chemical reactions.

Objective For Module Four

In completing this module, you will master the following objective:

- to balance chemical reactions

[†] A stoichiometric coefficient of 1 is usually omitted when writing a balanced chemical reaction; thus, the combustion of propane becomes

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Balanced Chemical Reactions and the Conservation of Mass

Before we learn how to balance a chemical reaction, it is worth reviewing the relationship between a balanced reaction and the conservation of mass. You may recall this paraphrase of one of Dalton's hypotheses for the existence of atoms:

> *In a chemical reaction the elements making up compounds rearrange to make new compounds. The atoms making up these compounds, however, are not destroyed, nor are new atoms created.*[†]

This statement that matter is conserved in a chemical reaction means that for every element present in the reactants, an equal amount of that element must be present in the products. When we write an unbalanced chemical reaction

$$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$$

we can show that mass is not conserved by comparing the number of atoms of each element on the reactant's and product's side of the arrow; thus

Element	Atoms in Reactants	Atoms in Products
C	3	1
H	8	2
O	2	3

As written, none of the elements is conserved so the reaction is unbalanced. The balanced chemical reaction

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

obeys the conservation of the mass.[‡]

Element	Atoms in Reactants	Atoms in Products
C	3	3
H	8	8
O	10	10

A balanced chemical reaction always obeys the conservation of mass.

[†] See Module 2 for a review of Dalton's hypotheses.

[‡] When counting atoms for a molecule with a stoichiometric coefficient, multiply the number of atoms in one molecule by the number of molecules. For example, $5CO_2$ has 5 x 1 = 5 carbon atoms and 5 x 2 = 10 oxygen atoms.

Balancing Chemical Reactions – No No's, Conventions, and Tips

Converting an unbalanced chemical reaction into one that is balanced is mostly a "trial and error" process. There are, however, some important things that you can't do, some common conventions, and some strategies that help simplify the process.

Things That You Can't Do When Balancing a Chemical Reaction. One of the most common mistakes when balancing a chemical reaction is to change the subscripts on compounds instead of changing the stoichiometric coefficients. For example, in the presence of a spark, gaseous mixtures of H_2 and O_2 react forming water as the only product. The unbalanced reaction based on this description, which is called a skeletal reaction, is

$$H_2 + O_2 \rightarrow H_2O$$

When balancing this reaction it is tempting to just add a subscripted 2 to the oxygen in the water molecule, giving

$$H_2 + O_2 \rightarrow H_2O_2$$

The problem with this is that H_2O_2 is the chemical formula for hydrogen peroxide, not water. Although this reaction is balanced, it is no longer the reaction of interest.

Another common mistake is to add new reactants or products to the reaction. For example, balancing the skeletal reaction $H_2 + O_2 \rightarrow H_2O$ by adding an oxygen atom as a second product

$$H_2 + O_2 \rightarrow H_2O + O$$

is incorrect because water is the reaction's only identified product. Although this reaction may take place under appropriate conditions (such as at high elevations in the atmosphere), it isn't the reaction with which we are working.

Common Conventions for Balanced Reactions. There are two common conventions for balanced reactions. First, because we cannot have a fraction of a molecule, the stoichiometric coefficients in a balanced reaction are usually written as integers. Although

$$H_2 + \tfrac{1}{2}O_2 \rightarrow H_2O$$

is a balanced reaction, it is more appropriate to multiply each stoichiometric coefficient by 2, obtaining

$$2H_2 + O_2 \rightarrow 2H_2O$$

Second, the stoichiometric coefficients should be reduced to the smallest whole numbers. For example, it is preferable to write the balanced reaction

$$2C_3H_8 + 10O_2 \rightarrow 6CO_2 + 8H_2O$$

as

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

by dividing each stoichiometric coefficient by 2.

Useful Tips for Balancing Chemical Reactions. Balancing a chemical reaction can be frustrating. The most common problem is discovering that balancing one element causes a previously balanced element to become unbalanced. This process can go on and on until you are ready to explode. The following three tips will help you avoid spontaneous combustion!

> *Tip #1 – Begin with elements that appear in only 1 reactant and 1 product, and end with those elements that appear in more than one reactant or product.*

The rationale for this tip is that it is easy to balance a reactant and product that are the only source of an element. In addition, once the stoichiometric ratio between the reactant and product is established, any change to the stoichiometric coefficient for one is easily transferred to the other. For example, consider the hypothetical skeletal reaction

$$A_2B + C_2 \rightarrow A + B_2 + BC$$

Following Tip #1, we first balance A and C because each appears in a single reactant and a single product

$$A_2B + C_2 \rightarrow 2A + B_2 + 2BC$$

We next balance element B, which appears in one reactant and two products. In doing so, we need to change the coefficient in front of A_2B from 1 to 4. Because the stoichiometry between A_2B and A has already been established at 1:2, we must adjust this to 4:8; thus, leaving the following balanced reaction.

$$4A_2B + C_2 \rightarrow 8A + B_2 + 2BC$$

> *Tip #2 – When balancing an element that appears in more than one reactant and one product, try to bring it into balance by adjusting the coefficient for a species that has not yet been assigned.*

The rationale for this tip is to avoid changing a coefficient that was adjusted earlier when bringing another element into balance. For example, consider the following hypothetical skeletal reaction

$$A_2B_8C_2 \rightarrow A + B_2 + BC$$

Both A and C appear in only a single reactant and a single product, so these are balanced first.

$$A_2B_8C_2 \rightarrow 2A + B_2 + 2BC$$

To balance B, we must choose between adjusting the coefficient for B_2 or BC. Because we have already adjusted the coefficient for BC in balancing C, any change to its coefficient will bring C out of balance. Instead, we adjust the coefficient for B_2, giving

$$A_2B_8C_2 \rightarrow 2A + 3B_2 + 2BC$$

> *Tip #3 – Whenever possible, balance the simplest compounds (pure elements or diatomic molecules) last.*

Because an element or diatomic molecule contains only one type of atom, any change to its coefficient cannot bring any other element out of balance. Furthermore, with a diatomic molecule, we can use a fractional coefficient of 1/2 to add a single atom. Of course, once balanced all coefficients are doubled to ensure that they are integers. For example, consider the following hypothetical skeletal reaction

$$AB_4C_3 + A_2D \rightarrow B_2D + CD + A_2$$

We begin by balancing B and C as each is present in a single reactant and product.

$$AB_4C_3 + A_2D \rightarrow 2B_2D + 3CD + A_2$$

Both A and D appear in more than one reactant or product. Because A appears by itself in the diatomic species A_2, it is easier to leave A for last; thus we balance D

$$AB_4C_3 + 5A_2D \rightarrow 2B_2D + 3CD + A_2$$

and then balance A

$$AB_4C_3 + 5A_2D \rightarrow 2B_2D + 3CD + {}^{11}\!/_{2}A_2$$

$$2AB_4C_3 + 10A_2D \rightarrow 4B_2D + 6CD + 11A_2$$

Balancing Chemical Reactions – Worked Examples

Having discussed some general procedures for balancing reactions, we are ready to work through some examples. In doing so, we will move from easy reactions to those that are more complex. Each example shows all reactants and products so that no knowledge about the underlying chemistry is necessary. Although we won't include them in our worked examples, you may find it helpful to use a table to keep track of atoms on the reactant and product side of the reaction (see page 3 for an example).

Example 1. Balance the decomposition reaction for potassium perchlorate, $KClO_4$.

$$KClO_4 \rightarrow KCl + O_2$$

Solution. The elements K and Cl are already balanced. To balance oxygen, we place a 2 before O_2 giving the final balanced reaction.

$$KClO_4 \rightarrow KCl + 2O_2$$

The first example is easy because only one element needs to be balanced and no adjustments to other coefficients are necessary. In the next example, only one element is out of balance, but bringing it into balance necessitates changing other coefficients.

Example 2. Balance the decomposition reaction for potassium chlorate, $KClO_3$.

$$KClO_3 \rightarrow KCl + O_2$$

Solution. Only oxygen is not balanced, which we balance by adding a coefficient of 3/2 before O_2.

$$KClO_3 \rightarrow KCl + {}^{3}\!/_{2}\,O_2$$

We then multiply all the coefficients by 2 to give the final balanced reaction.

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

For most reactions, two or more elements in the skeletal reaction are out of balance. The next set of examples provides good illustrations of balancing such reactions.

Example 3. Balance the combustion reaction for methane, CH_4.

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

Solution. First we balance H, which is present in only one reactant and one product.

$$CH_4 + O_2 \rightarrow CO_2 + 2H_2O$$

Everything is now balanced except oxygen, for which there are 2 on the reactant's side and 4 on the product's side. Adding a 2 before the O_2 provides the balanced reaction.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Example 4. Balance the combustion reaction for propane, C_3H_8.

$$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$$

Solution. We begin by balancing C and H because each is present in a single reactant and a single product.

$$C_3H_8 + O_2 \rightarrow 3CO_2 + 4H_2O$$

This leaves us with 2 O atoms on the reactant's side and 10 O atoms on the product's side; thus, we add a 5 before O_2 to give the balanced reaction.

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Example 5. Balance the combustion reaction for benzoic acid, $C_7H_6O_2$.

$$C_7H_6O_2 + O_2 \rightarrow CO_2 + H_2O$$

Solution. As in the previous example, we begin by balancing C and H.

$$C_7H_6O_2 + O_2 \rightarrow 7CO_2 + 3H_2O$$

This leaves 4 O atoms on the reactant's side and 17 O atoms on the product's side. To avoid unbalancing C and H, we balance O by adding a coefficient of 15/2 before O_2; thus

$$C_7H_6O_2 + {}^{15}\!/_{2}\,O_2 \rightarrow 7CO_2 + 3H_2O$$

$$2C_7H_6O_2 + 15O_2 \rightarrow 14CO_2 + 6H_2O$$

Example 6. Balance the following oxidation reaction for the mineral pyrite, FeS_2.

$$FeS_2 + O_2 \rightarrow Fe_2O_3 + SO_2$$

Solution. We begin by balancing Fe

$$2FeS_2 + O_2 \rightarrow Fe_2O_3 + SO_2$$

and then S

$$2FeS_2 + O_2 \rightarrow Fe_2O_3 + 4SO_2$$

Finally, we balance O by adjusting the coefficient for O_2

$$2FeS_2 + {}^{11}\!/_{2}\,O_2 \rightarrow Fe_2O_3 + 4SO_2$$

$$4FeS_2 + 11O_2 \rightarrow 2Fe_2O_3 + 8SO_2$$

Example 7. Balance the following reaction for dissolving silver.

$$Ag + H_2SO_4 \rightarrow Ag_2SO_4 + SO_2 + H_2O$$

Solution. Balancing Ag is easy, leaving us with

$$2Ag + H_2SO_4 \rightarrow Ag_2SO_4 + SO_2 + H_2O$$

Next, we balance S, by placing a 2 before H_2SO_4.

$$2Ag + 2H_2SO_4 \rightarrow Ag_2SO_4 + SO_2 + H_2O$$

Finally, we place a 2 before H_2O to balance H and O.

$$2Ag + 2H_2SO_4 \rightarrow Ag_2SO_4 + SO_2 + 2H_2O$$

Example 8. Balance the following reaction for dissolving gold.

$$Au + HNO_3 + HCl \rightarrow HAuCl_4 + NO + H_2O$$

Solution. We begin by balancing the chlorine.

$$Au + HNO_3 + 4HCl \rightarrow HAuCl_4 + NO + H_2O$$

Next, we balance H, adding a coefficient of 2 before H_2O, giving a balanced reaction of

$$Au + HNO_3 + 4HCl \rightarrow HAuCl_4 + NO + 2H_2O$$

Now for more of a challenge! Note how using the concept of a structural unit, as opposed to working only with elements, helps simplify the process of balancing the reaction.

Example 9. Balance the following reaction.

$$H_3PO_4 + (NH_4)_2MoO_4 + HNO_3 \rightarrow (NH_4)_3PO_4{\bullet}12MoO_3 + NH_4NO_3 + H_2O$$

Solution. We begin by balancing Mo, which shows up in one reactant and one product, placing a 12 before $(NH_4)_2MoO_4$

$$H_3PO_4 + 12(NH_4)_2MoO_4 + HNO_3 \rightarrow (NH_4)_3PO{\bullet}12MoO_3 + NH_4NO_3 + H_2O$$

Next, we balance the structural unit NH_4 (actually the ammonium ion, NH_4^+), which appears in $(NH_4)_2MoO_4$, $(NH_4)_3PO_4{\bullet}12MoO_3$, and NH_4NO_3. There are 24 NH_4 units on the reactant's side and 4 on the product's side. Following the advice of Tip #2, we adjust the coefficient for NH_4NO_3 instead of $(NH_4)_3PO_4{\bullet}12MoO_3$ to avoid throwing Mo out of balance.

$$H_3PO_4 + 12(NH_4)_2MoO_4 + HNO_3 \rightarrow (NH_4)_3PO_4{\bullet}12MoO_3 + 21NH_4NO_3 + H_2O$$

Next, we balance the structural unit NO_3 (actually the nitrate ion, NO_3^-), which appears in HNO_3 and NH_4NO_3.

$$H_3PO_4 + 12(NH_4)_2MoO_4 + 21HNO_3 \rightarrow (NH_4)_3PO_4{\bullet}12MoO_3 + 21NH_4NO_3 + H_2O$$

Finally, we balance H, giving

$$H_3PO_4 + 12(NH_4)_2MoO_4 + 21HNO_3 \rightarrow (NH_4)_3PO_4{\bullet}12MoO_3 + 21NH_4NO_3 + 12H_2O$$

Balancing Chemical Reactions – Including Ions

Many of the examples use reactions involving inorganic compounds. Such reactions often include ions that never undergo a change in chemistry. These ions are called spectator ions and are not actually a part of the reaction. For example, soluble salts of the silver ion, Ag^+, will form solid AgCl, which is called a precipitate, when reacted with any soluble salt containing the chloride ion, Cl^-. The following balanced reactions

$$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$$

$$2AgNO_3 + CaCl_2 \rightarrow 2AgCl + Ca(NO_3)_2$$

can be written simply as

$$Ag^+ + Cl^- \rightarrow AgCl$$

by ignoring the spectator ions (NO_3^-, Ca^{2+}). Although you aren't expected in this course to recognize which ions are spectators, you should be able to balance a reaction including ions.

One important caution: a balanced reaction including ions must have the same net charge on each side of the reaction's arrow. For example

$$Ag^+ + Cu \rightarrow Cu^{2+} + Ag$$

isn't balanced because the reactant's side has a net charge of +1 from Ag^+, whereas the product's side has a net charge of +2 from Ca^{+2}. The correct balanced reaction is

$$2Ag^+ + Cu \rightarrow Cu^{2+} + 2Ag$$

Example 10. Balance the following reaction.

$$S_2O_8^{2-} + H_2O + Mn^{2+} \rightarrow MnO_2 + H^+ + SO_4^{2-}$$

Solution. We begin by balance S, giving

$$S_2O_8^{2-} + H_2O + Mn^{2+} \rightarrow MnO_2 + H^+ + 2SO_4^{2-}$$

Next, we balance O by placing a 2 before the H_2O

$$S_2O_8^{2-} + 2H_2O + Mn^{2+} \rightarrow MnO_2 + H^+ + 2SO_4^{2-}$$

saving H^+ for last as it is easy to balance a single element.

$$S_2O_8^{2-} + 2H_2O + Mn^{2+} \rightarrow MnO_2 + 4H^+ + 2SO_4^{2-}$$

Note that each side of the reaction has a net charge of zero.

Balancing Chemical Reactions – Complications

Occasionally a reaction proves particularly difficult to balance. As an exercise (and to appreciate the challenge some reactions present), try balancing the following reaction. Don't spend more than about five minutes on this exercise.

$$Cu + NO_3^- + H_3O^+ \rightarrow Cu^{2+} + NO + H_2O$$

Were you able to balance the reaction? Don't be surprised (or disappointed) if your answer is no.

What makes this reaction difficult to balance is the presence of oxygen in two reactants, NO_3^- and H_3O^+, and in two products, NO and H_2O. Our simple rules for balancing reactions are less useful in this case. You can reach the correct answer, which is

$$3Cu + 2NO_3^- + 8H_3O^+ \rightarrow 3Cu^{2+} + 2NO + 12H_2O$$

by a combination of trial-and-error and a little logic, but the time and effort expended can be significant.

As difficult as the above reaction may be to balance, eventually you can, with some effort and patience, arrive at a correctly balanced reaction. Unfortunately, this is not always the case. Consider, for example, the following unbalanced reaction

$$MnO_4^- + H_2O_2 + H_3O^+ \rightarrow Mn^{2+} + O_2 + H_2O$$

Here are two solutions that meet our criteria for a balanced reaction, although both solutions actually are chemically incorrect!

$$2MnO_4^- + H_2O_2 + 6H_3O^+ \rightarrow 2Mn^{2+} + 3O_2 + 10H_2O$$

$$2MnO_4^- + 3H_2O_2 + 6H_3O^+ \rightarrow 2Mn^{2+} + 4O_2 + 12H_2O$$

We'll consider how to balance these types of reactions in the next section.

Balancing Chemical Reactions – An Alternative Approach

In the previous section we showed two reactions that appear to be "balanced" and yet are chemically incorrect. How can this be true? The answer to this question requires us to see that the reaction we are trying to balance involves a transfer of electrons from one reactant to another reactant. We call such reactions oxidation/reduction or redox reactions.

Oxidation States. To understand what happens during the (unbalanced) reaction

$$Cu + NO_3^- + H_3O^+ \rightarrow Cu^{2+} + NO + H_2O$$

we must introduce the concept of an oxidation state.[‡] An oxidation state is a means for keeping track of electrons in a chemical reaction. A few simple rules will help us assign oxidation states in this reaction:

‡ Although we introduce the concept of oxidation states here to help us understand the logic behind this alternative approach for balancing redox reactions, you can balance this or any redox reaction without knowing the oxidation states of elements in the reaction; in fact, you can use this alternative approach to balancing reactions that do not involve changes in oxidation state (although there is no need to do so).

Rule #1. The oxidation state of any element in its elemental form is zero; thus, the oxidation state for Cu is zero.

Rule #2. The oxidation state for a cation or anion consisting of a single element is the same as the ion's charge; thus, the oxidation state of copper in Cu^{2+} is +2.

Rule #3. In compounds and ions, hydrogen <u>always</u> has an oxidation state of +1 when bound to a non-metal, such as oxygen.

Rule #4. In compounds and ions, oxygen usually has an oxidation state of -2.

Rule #5. The algebraic sum of oxidation states for the elements in a polyatomic compound or ion must equal the compound's total charge; thus

for NO_3^-: 3 × (oxidation state of O) + oxidation state of N = -1

for NO: oxidation state of N + oxidation state of O = 0

for H_3O^+: 3 × (oxidation state of H) + oxidation state of O = +1

for H_2O: 2 × (oxidation state of H) + oxidation state of O = 0

Applying these rules to the compounds and ions in the (unbalanced) reaction

$$Cu + NO_3^- + H_3O^+ \rightarrow Cu^{2+} + NO + H_2O$$

we find the following oxidation states:

Copper: oxidation states of zero in Cu and +2 in Cu^{2+}

Oxygen: an oxidation state of -2 in NO_3^-, H_3O^+, NO and H_2O

Hydrogen: an oxidation state of +1 in H_3O^+ and H_2O

Nitrogen: an oxidation state of +5 in NO_3^- and +2 in NO

Oxidation and Reduction. An element experiencing an increase in its oxidation state loses electrons and is said to undergo oxidation. For example, in the (unbalanced) reaction

$$Cu + NO_3^- + H_3O^+ \rightarrow Cu^{2+} + NO + H_2O$$

the copper in Cu is oxidized when forming Cu^{2+} (a change in oxidation state from zero to +2). When an element gains electrons it experiences a decrease in its oxidation state and is said to be reduced. Thus, in the reaction shown above, the nitrogen in NO_3^- is reduced when forming NO (a change in oxidation state from +5 to +2).

Redox Reactions. Reducing the nitrogen in NO_3^- to NO requires adding electrons. The source of these electrons is the oxidation of copper from Cu to Cu^{2+}. Thus, any reaction in which one reactant experiences reduction must have another reactant that undergoes oxidation. We call such reactions oxidation/reduction or redox reaction.

The Alternative Approach to Balancing Reactions. Because a balanced redox reaction does not include electrons as reactants or products, all electrons released by the species undergoing oxidation must be consumed by the species undergoing reduction. This is the key to balancing redox reactions. Here is our general approach.

$$Cu + NO_3^- + H_3O^+ \rightarrow Cu^{2+} + NO + H_2O$$

Step 1. Eliminate any H_2O, H_3O^+ and OH^- present in the unbalanced reaction. Because the reactions we will consider always occur in water, we can add these species back in at any time. This leaves us with

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

Step 2. Separate the reaction into two parts representing the oxidation and reduction processes. Note – even if you don't know which species are undergoing oxidation and reduction, the two reactions should be obvious. This leave us with

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

Step 3. Balance all elements in each reaction *except for oxygen and hydrogen*. In this case the copper and nitrogen already are balanced so no adjustments are needed.

Step 4. Balance the oxygen in each reaction by adding water, H_2O. Since there are three oxygens in NO_3^- and only one oxygen in NO, we add two molecules of H_2O to the products of the second reaction. This leaves us with

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

Step 5. Balance the hydrogen in each reaction by adding a combination of H_3O^+ and H_2O. Note that an ion of H_3O^+ has one more hydrogen than a molecule of H_2O; thus, adding an equal number of H_3O^+ ions and H_2O molecules to opposite sides of a reaction has the effect of increasing the number of hydrogens on the side of the reaction receiving the H_3O^+ ions by the number of H_3O^+ ions added. For example, since there are four hydrogens in the products and none in the reactants, we need to add the equivalent of four hydrogens to the reactants. We accomplish this by adding four H_3O^+ ions to the reactants and four H_2O molecules to the products (a net gain of four hydrogens by the reactants). This leaves us with

$$Cu \rightarrow Cu^{2+} \qquad\qquad NO_3^- + 4H_3O^+ \rightarrow NO + 6H_2O$$

Note – for basic solutions we add H_2O and OH^- instead of H_3O^+ and H_2O. For example, if the above reaction were to occur in a basic solution, we would add four H_2O molecules to the reactants and four OH^- ions to the products (a net increase gain of four hydrogens by the reactants)

$$NO_3^- + 4H_2O \rightarrow NO + 2H_2O + 4OH^-$$

which simplifies to

$$NO_3^- + 2H_2O \rightarrow NO + 4OH^-$$

Step 6. Balance the charge by adding electrons (e^-). Note that the electrons must appear as a product in one reaction and as a reactant in the other reaction. Because the first reaction has a net charge of zero on the reactant side and a net charge of +2 on the product side, we add two electrons to the products. For the second reaction we need to add three electrons to the reactants to balance out the charge. This leaves us with

$$Cu \rightarrow Cu^{2+} + 2e^- \qquad\qquad NO_3^- + 4H_3O^+ + 3e^- \rightarrow NO + 6H_2O$$

Note – when using this approach to balance a non-redox reaction, the charge will be balanced without the need to add electrons.

Step 7. Before combining the two reactions the number of electrons must be the same so that no electrons will remain in the final balanced reaction. To accomplish this we multiply each coefficient in the first reaction by three and each coefficient in the second reaction by two

$3(Cu \rightarrow Cu^{2+} + 2e^-)$	$2(NO_3^- + 4H_3O^+ + 3e^- \rightarrow NO + 6H_2O)$

leaving us with six electrons in each reaction

$$3Cu \rightarrow 3Cu^{2+} + 6e^- \qquad\qquad 2NO_3^- + 8H_3O^+ + 6e^- \rightarrow 2NO + 12H_2O$$

Step 8. Finally, add the two reactions together and simplify as needed. This leaves us with a balanced reaction with no left over electrons.

$$3Cu + 2NO_3^- + 8H_3O^+ \rightarrow 3Cu^{2+} + 2NO + 12H_2O$$

Example 11. Find the correct balanced reaction for

$$MnO_4^- + H_2O_2 + H_3O^+ \rightarrow Mn^{2+} + O_2 + H_2O$$

Solution. Using our alternative approach we first eliminate the H_3O^+ and H_2O

$$MnO_4^- + H_2O_2 \rightarrow Mn^{2+} + O_2$$

Next, we split the reaction into two parts, one involving manganese and the other involving oxygen

$$MnO_4^- \rightarrow Mn^{2+} \qquad H_2O_2 \rightarrow O_2$$

Since the manganese already is balanced, we next balance oxygen by adding H_2O

$$MnO_4^- \rightarrow Mn^{2+} + 4H_2O \qquad H_2O_2 \rightarrow O_2$$

To balance the hydrogen in the reaction on the left, where we need to add eight hydrogens to the reactants, we add eight H_3O^+ ions to the reactants and eight H_2O molecules to the products. To balance the hydrogen in the reaction on the right, where we need to add two hydrogens to the products, we add two molecules of H_2O to the reactants and two molecules of H_3O^+ to the products; thus

$$MnO_4^- + 8H_3O^+ \rightarrow Mn^{2+} + 12H_2O \qquad H_2O_2 + 2H_2O \rightarrow O_2 + 2H_3O^+$$

Next we balance charge by adding electrons

$$MnO_4^- + 8H_3O^+ + 5e^- \rightarrow Mn^{2+} + 12H_2O \qquad H_2O_2 + 2H_2O \rightarrow O_2 + 2H_3O^+ + 2e^-$$

and adjust the coefficients so that each reaction involves 10 electrons

$$2MnO_4^- + 16H_3O^+ + 10e^- \rightarrow 2Mn^{2+} + 24H_2O \qquad 5H_2O_2 + 10H_2O \rightarrow 5O_2 + 10H_3O^+ + 10e^-$$

Adding the reactions together and simplifying gives the balanced reaction as

$$2MnO_4^- + 5H_2O_2 + 6H_3O^+ \rightarrow 2Mn^{2+} + 5O_2 + 14H_2O$$

Here is an example that involves a reaction in a basic solution.

Example 12. Balance the following reaction, which occurs in basic solutions.

$$CuO + NH_3 \rightarrow Cu + N_2$$

Solution. Dividing the reaction into two parts gives

$$CuO \rightarrow Cu \qquad NH_3 \rightarrow N_2$$

Next, we balance the nitrogen in the second reaction, giving

$$CuO \rightarrow Cu \qquad 2NH_3 \rightarrow N_2$$

To balance the oxygen in the first reaction we add one molecule of H_2O

$$CuO \rightarrow Cu + H_2O \qquad 2NH_3 \rightarrow N_2$$

Because the solution is basic, we balance hydrogen by adding H_2O and OH^-. Because the first reaction has two hydrogens on the product's side we add two units of H_2O to the reactants and two units of OH^- to the products, giving a net increase of two hydrogens to the reactant's side of the reaction. Using the same logic, we add six units of H_2O to the products of the second reaction and six units of OH^- to the reactants; thus

$$CuO + 2H_2O \rightarrow Cu + H_2O + 2OH^- \qquad 2NH_3 + 6OH^- \rightarrow N_2 + 6H_2O$$

Simplifying the first reaction by removing one unit of H_2O from both sides leave us with

$$CuO + H_2O \rightarrow Cu + 2OH^- \qquad 2NH_3 + 6OH^- \rightarrow N_2 + 6H_2O$$

Next we balance charge by adding electrons, giving

$$CuO + H_2O + 2e^- \rightarrow Cu + 2OH^- \qquad 2NH_3 + 6OH^- \rightarrow N_2 + 6H_2O + 6e^-$$

Multiplying the coefficients of the first reaction by three

$$3CuO + 3H_2O + 6e^- \rightarrow 3Cu + 6OH^- \qquad 2NH_3 + 6OH^- \rightarrow N_2 + 6H_2O + 6e^-$$

gives each reaction the same number of electrons. Adding the reactions together and simplifying gives the final balanced reaction as

$$3CuO + 2NH_3 \rightarrow 3Cu + N_2 + 3H_2O$$

Here is an unusual example of a reaction to balance in that it has only a single identified product. Note, however, that the alternative approach still works.

Example 13. Balance the following reaction, assuming that the solution is acidic.

$$HIO_3 + HI \rightarrow I_2$$

Solution. As with previous problems, we begin by dividing the reaction into two parts. Although the reaction shows only one product, I_2, both reactants include iodine; thus, they both must be converted into I_2. This leaves us with the following two reactions

$$HIO_3 \rightarrow I_2 \qquad HI \rightarrow I_2$$

Balancing iodine in both reactions leave us with

$$2HIO_3 \rightarrow I_2 \qquad 2HI \rightarrow I_2$$

Next we balance oxygen by adding H_2O

$$2HIO_3 \rightarrow I_2 + 6H_2O \qquad 2HI \rightarrow I_2$$

Because the reaction on the left has two hydrogens on the product side and 12 hydrogens on the reactant side, we need to add an additional 10 hydrogens to the products. We accomplish this by adding 10 H_3O^+ ions to the reactants and 10 additional H_2O molecules to the products (giving the products a total of 16 H_2O molecules). Balancing hydrogen for the reaction on the right requires adding two hydrogens to the products, which we accomplish by adding two molecules of H_2O to the reactants and two H_3O^+ ions to the products. This leaves use with

$$2HIO_3 + 10H_3O^+ \rightarrow I_2 + 16H_2O \qquad 2HI + 2H_2O \rightarrow I_2 + 2H_3O^+$$

Adding electrons to balance charge

$$2HIO_3 + 10H_3O^+ + 10e^- \rightarrow I_2 + 16H_2O \qquad 2HI + 2H_2O \rightarrow I_2 + 2H_3O^+ + 2e^-$$

and multiplying the coefficients for the second reaction by five leaves both reactions with 10 electrons; thus

$$2HIO_3 + 10H_3O^+ + 10e^- \rightarrow I_2 + 16H_2O \qquad 10HI + 10H_2O \rightarrow 5I_2 + 10H_3O^+ + 10e^-$$

Combining the reactions and simplifying gives the final balanced reaction

$$HIO_3 + 5HI \rightarrow 3I_2 + 3H_2O$$

Practice Problems

The following problems provide practice in meeting this module's objectives. Answers are provided on the last page. Be sure to seek assistance if you experience difficulty with any of these problems. When you are ready, schedule an appointment for the module's exam.

1. When I took high school chemistry we did an experiment where we heated a sample of ammonium dichromate, $(NH_4)_2Cr_2O_7$, which proceeded to "erupt" like a volcano, spewing out gases and leaving behind a residue of chromium oxide, Cr_2O_3. Balance the skeletal reaction

$$(NH_4)_2Cr_2O_7 \rightarrow N_2 + Cr_2O_3 + H_2O$$

2. There are relatively few reactions at room temperature that involve only solid reactants. One such reaction occurs when shaking together barium hydroxide octahydrate, $Ba(OH)_2 \bullet 8H_2O$, and ammonium thiocyanate, NH_4SCN. Balance the skeletal reaction

$$Ba(OH)_2 \bullet 8H_2O + NH_4SCN \rightarrow Ba(SCN)_2 + H_2O + NH_3$$

3. Balance the following skeletal reaction for the combustion of sucrose

$$C_{12}H_{22}O_{11} + O_2 \rightarrow CO_2 + H_2O$$

4. Balance the following skeletal reaction for the combustion of ethanol

$$C_2H_6O + O_2 \rightarrow CO_2 + H_2O$$

5. Balance the following skeletal reaction for the combustion of benzene

$$C_6H_6 + O_2 \rightarrow CO_2 + H_2O$$

6. Aspartame, $C_{14}H_{18}N_2O_5$, was discovered by a graduate of DePauw. Balance the following skeletal reaction for its combustion

$$C_{14}H_{18}N_2O_5 + O_2 \rightarrow CO_2 + H_2O + N_2$$

7. Hydrogen cyanide, HCN, which is a nasty, poisonous gas, is produced industrially by reacting together ammonia, oxygen, and methane. Balance the following skeletal reaction for its synthesis

$$NH_3 + O_2 + CH_4 \rightarrow HCN + H_2O$$

8. Nitric acid, HNO_3, is produced by the Ostwald process, which consists of the following three unbalanced reactions; balance each.

$$NH_3 + O_2 \rightarrow NO + H_2O$$

$$NO + O_2 \rightarrow NO_2$$

$$NO_2 + H_2O \rightarrow HNO_3 + NO$$

9. Balance the following skeletal reaction of a strong acid, H^+, with calcium bicarbonate, $Ca(HCO_3)_2$

$$H^+ + Ca(HCO_3)_2 \rightarrow Ca^{2+} + H_2O + CO_2$$

10. Here is a more complicated problem to balance

$$K_4Fe(CN)_6 + H_2SO_4 + H_2O \rightarrow K_2SO_4 + FeSO_4 + (NH_4)_2SO_4 + CO$$

11. Sodium metal, Na, reacts with chlorine gas, Cl_2, to give sodium chloride, NaCl. Write a balanced chemical equation for this reaction.

12. Iron, Fe, forms a variety of iron oxides upon reacting with oxygen. Write balanced reactions showing the formation of each of the following: FeO, Fe_2O_3, and Fe_3O_4. In each case, the iron oxide is the reaction's only product.

13. Upon heating, lead nitrate, $Pb(NO_3)_2$, explodes, forming lead oxide, PbO, nitrogen dioxide, NO_2, and oxygen, O_2, as products. Write a balanced chemical equation for this reaction.

14. Balance the following reaction between chromate, CrO_4^-, and manganese ion, Mn^{2+}. You may assume that the reaction occurs in an acidic solution.

$$CrO_4^- + Mn^{2+} \rightarrow Cr^{3+} + MnO_4^-$$

15. Balance the following reaction between oxalic acid, $C_2H_2O_4$, and permanganate, MnO_4^-. You may assume that the reaction occur in an acidic solution.

$$C_2H_2O_4 + MnO_4^- \rightarrow CO_2 + Mn^{2+}$$

16. Example 11 shows the balanced reaction between permanganate, MnO_4^-, and hydrogen peroxide, H_2O_2, in an acidic solution. In a basic solution the permanganate reduces to MnO_2 instead of Mn^{2+}. What is the complete balanced reaction?

17. Balance the following reaction between ammonia, NH_3, and hypochlorite, OCl^-, forming hydrazine, N_2H_4, and chloride, Cl^-. You may assume that the reaction occurs in a basic solution.

$$NH_3 + OCl^- \rightarrow N_2H_4 + Cl^-$$

18. Balance the following reaction in which nitrous acid, HNO_2, reacts with itself (what is commonly called a disproportionation reaction). You may assume that the reaction occurs in an acidic solution.

$$HNO_2 \rightarrow NO_3^- + NO$$

Hint: Begin by writing two reactions, both of which have HNO_2 as a reactant.

19. Balance the following reaction between sulfur dioxide, SO_2, and hydrogen sulfide, H_2S. You may assume that the reaction occurs in an acidic solution.

$$SO_2 + H_2S \rightarrow S$$

Answers to Practice Problems

1. $(NH_4)_2Cr_2O_7 \rightarrow N_2 + Cr_2O_3 + 4H_2O$

2. $Ba(OH)_2{\bullet}8H_2O + 2NH_4SCN \rightarrow Ba(SCN)_2 + 10H_2O + 2NH_3$

3. $C_{12}H_{22}O_{11} + 12O_2 \rightarrow 12CO_2 + 11H_2O$

4. $C_2H_6O + 3O_2 \rightarrow 2CO_2 + 3H_2O$

5. $2C_6H_6 + 15O_2 \rightarrow 12CO_2 + 6H_2O$

6. $C_{14}H_{18}N_2O_5 + 16O_2 \rightarrow 14CO_2 + 9H_2O + N_2$

7. $2NH_3 + 3O_2 + 2CH_4 \rightarrow 2HCN + 6H_2O$

8. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

   $2NO + O_2 \rightarrow 2NO_2$

   $3NO_2 + H_2O \rightarrow 2HNO_3 + NO$

9. $2H^+ + Ca(HCO_3)_2 \rightarrow Ca^{2+} + 2H_2O + 2CO_2$

10. $K_4Fe(CN)_6 + 6H_2SO_4 + 6H_2O \rightarrow 2K_2SO_4 + FeSO_4 + 3(NH_4)_2SO_4 + 6CO$

11. $2Na + Cl_2 \rightarrow 2NaCl$

12. $2Fe + O_2 \rightarrow 2FeO$

    $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

    $3Fe + 2O_2 \rightarrow Fe_3O_4$

13. $2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$

14. $5CrO_4^- + 4Mn^{2+} + 8H_3O^+ \rightarrow 5Cr^{3+} + 4MnO_4^- + 12H_2O$

15. $5C_2H_2O_4 + 2MnO_4^- +6H_3O^+ \rightarrow 10CO_2 + 2Mn^{2+} + 14H_2O$

16. $2MnO_4^- + 3H_2O_2 \rightarrow 2MnO_2 + 3O_2 + 2OH^- + 2H_2O$

17. $2NH_3 + OCl^- \rightarrow N_2H_4 + Cl^- + H_2O$

18. $3HNO_2 \rightarrow NO_3^- + 2NO + H_3O^+$

19. $SO_2 + 2H_2S \rightarrow 3S + 2H_2O$