Chemical Reactions

Physical changes do not form new substances. With physical changes, the chemical or atomic composition of a molecule stays the same. For example, boiling water turns it from a liquid to a gas, but the molecule is water, H_2O.

Chemical changes form new substances. Many everyday processes involve chemical change. Some examples include life (respiration), driving a car (burning gasoline), rusting, cooking, and rotting garbage.

Chemical change can be difficult to recognize from laboratory (or physical) evidence. There are four types of physical evidence that indicate chemical change occurred: gas releases, solid forms (precipitate), permanent color change, and heat or light is released (or absorbed).

Evidence is not proof of chemical change because physical changes can give similar evidence. For example, boiling water produces a gas. Sugar crystallizes, which means going from aqueous solution to a solid. Ice forms and changes color from colorless to white. Boiling water absorbs heat. Freezing water releases heat. These are all examples physical changes.

Use of the word "precipitate" or "precipitation" means that a chemical change occurred that produced a solid. Use of the term "crystallization" means a physical change. A solid appears with both of these changes, one chemical and one physical. Use of the words "oxidation" and "reduction" also mean chemical change, and you will learn more about them in Week 10.

When writing a chemical reaction, reactant(s) appear on the left side, and product(s) appear on the right side of the reaction arrow, →. In this example reaction, A and B are reactants, and C and D are products: $A + B \rightarrow C + D$.

State symbols indicate the physical form of the reactants and products: (s) for solid, (l) for liquid, (g) for gas, and (aq) for and aqueous solution (meaning dissolved in water). A catalyst speeds up a reaction, and when one is used, it appears above the reaction arrow. Heat, Δ, is often a catalyst. Elements or compounds (written as chemical symbols) can be catalysts. In this example reaction, liquid reactant A and solid reactant B are heated to form, gaseous product C and aqueous solution D: $A(l) + B(s) \xrightarrow{\Delta} C(g) + D(aq)$.

The letters NR mean that no reaction occurred. In the previous example, this would be: $A(l) + B(s) \xrightarrow{\Delta} NR$.

Combining this material with that from Week 6, we can describe the chemicals in a chemical reaction: $S(s) + O_2(g) \rightarrow SO_2(l)$. This reaction reads: solid sulfur reacts with oxygen gas to form liquid sulfur dioxide.

Seven elements always appear as diatomic (two atom) molecules in chemical reactions. These elements are hydrogen, oxygen, nitrogen, chlorine, bromine, iodine, and fluorine. The chemical symbols of the element can spell the word, HONClBrIF (prounounced huncl-brif). Note that oxygen gas in the previous chemical reaction appears as O_2, not as O.

When writing chemical equations, they must have the same numbers of atoms of each type on each side. The process of getting the atom count the same is called **balancing**. When the atom count for each type of atom is the same on the left and right sides, this is called a balanced chemical reaction. We balance chemical reactions using coefficients, numbers written before the chemical in a chemical reaction. (We never change a subscript to balance a chemical reaction. Subscripts give the correct chemical formula, which we learned in Week 6.) A **coefficient** multiplies all the atoms that come after it. For example, $3P_2O_5$ means a total of six phosphorous atoms (3x2) and 15 oxygen atoms (3x5). When balancing a chemical reaction, *never* change the 2 and 5 subscripts.

Guidelines to balance a chemical reaction
1) Make sure that the chemical formulas are written correctly. (If any chemical formulas are wrong, balancing may be impossible. The skill to write a correct formula comes from Week 6)

2) *Never* change the subscripts after a chemical formula is correctly written.

3) Balance each atom type in a chemical formula starting with the atoms in the most complex formula.

4) If they appear on both sides of a chemical reaction, balance polyatomic ions (these are given on the formulas page) as a single unit. If a polyatomic ion changes in a chemical reaction, balance the atoms singly as given in Step 3.

Example: $2HC_2H_3O_2(aq) + Na_2CO_3(s) \rightarrow 2NaC_2H_3O_2(aq) + CO_2(g) + H_2O(l)$. In this example, balance the acetate anion, $C_2H_3O_2^-$, as a single unit (2 acetates on the left, 2 acetates on the right). The carbonate anion, CO_3^{2-}, does not appear on the right side, so balance the C and O atoms separately.

5) Coefficients must be whole numbers. If you get a fraction, multiply all coefficients by the denominator of the fraction to clear the fraction.

Example: The coefficients in this reaction, $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$, must be multiplied by 2 to get: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$. (Note: oxygen is only diatomic as the element, and not after it forms a compound, H_2O).

6) After balancing the equation, check that the atom count is the same on both sides. If it is not, either continue (if that makes sense) or start over from the beginning.

7) Check that you have the smallest whole number ratio of coefficients. If you can divide all coefficients by a common factor, do so.

Example: The coefficients in this reaction, $4H_2(g) + 2O_2(g) \rightarrow 4H_2O(l)$, must be divided by 2 to get: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$.

Example for balancing a chemical equation:

$Al_2(SO_4)_3(aq) + Ca(NO_3)_2(aq) \rightarrow Al(NO_3)_3(aq) + CaSO_4(s)$.

Guideline 1) All compounds are correctly written. Continue.
Guideline 2) *Never* change the subscripts after this point. Continue.
Guideline 3 and 4) The aluminum atom count differs. Two Al atoms are on the left, and one is on the right.

__2__ Al __1__

Place a 2 coefficient before $Al(NO_3)_3$ on the right side, giving:

$Al_2(SO_4)_3(aq) + Ca(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + CaSO_4(s)$.

__2__ Al __2__

The sulfate (polyatomic ion) count is different with 3 on the left side and 1 on the right side.

__3__ SO_4 __1__

Place a 3 coefficient before $CaSO_4$ on the right side, giving:

$Al_2(SO_4)_3(aq) + Ca(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3CaSO_4(s)$

__2__ Al __2__
__3__ SO_4 __3__

The calcium atom count is different with one Ca atom on the left and 3 on the right.

__1__ Ca __3__

Place a 3 coefficient before $Ca(NO_3)_2$ on the left side, giving:

$Al_2(SO_4)_3(aq) + 3Ca(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3CaSO_4(s)$

__2__ Al __2__
__3__ SO_4 __3__
__3__ Ca __3__

The nitrate (polyatomic ion) count is balanced with 6 on each side, so we are finished.

$Al_2(SO_4)_3(aq) + 3Ca(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3CaSO_4(s)$

__2__ Al __2__
__3__ SO_4 __3__
__3__ Ca __3__
__6__ NO_3 __6__

Guideline 5) No coefficients are fractions. Continue.
Guideline 6) Atom counts are: 2 aluminum atoms left and 2 aluminum atoms right, 3 sulfate ions left and 3 sulfate ions right, 3 calcium atoms left and 3 calcium atoms right, and 6 nitrate atoms left and 6 nitrate atoms right. Continue.
Guideline 7) The coefficients 1, 2 and 3 are not reducible, Done.

For more practice, see the separate primers on balancing reactions. Practice balancing the later reactions in this lecture by removing the coefficients and then balancing to reach the reactions written below.

We can place chemical reactions into five categories:
1) Combination Reactions
2) Decomposition Reactions
3) Single-Replacement Reactions
4) Double-Replacement Reactions
5) Neutralization Reactions

We will look at three types of **combination reactions**, also called **synthesis reactions**.
1) The reaction of a metal with oxygen
Example: $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$ which reads - iron metal reacts with oxygen gas to give iron(III) oxide).
Example: $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$ which reads – sodium metal reacts with oxygen gas to give sodium oxide.

2) The reaction of a nonmetal with oxygen
Example: $S(s) + O_2(g) \rightarrow SO_2(g)$ which reads – solid sulfur reacts with oxygen gas to give gaseous sulfur dioxide.
Example: $P_4(s) + 3O_2(g) \rightarrow 2P_2O_3(a)$ which reads – solid phosphorous reacts with oxygen gas to give solid diphosphorous trioxide.

3) The reaction of a metal and a nonmetal
Example: $2Na(s) + F_2(g) \rightarrow 2NaF(s)$ which reads – solid sodium metal reacts with fluorine gas to give sodium fluoride salt. (Note: if the reaction involves a transition metal, the formula of the salt is not always predictable, whereas with a main group metal, the formula is always predictable – Week 6)

A **decomposition reaction** breaks a compound down into two or more components, and it usually requires heat or light to make it happen.

We will look at three types of decomposition reactions here.
1) Ionic compounds that are oxides often decompose into metal plus oxygen gas.
Example: $2HgO(s) \xrightarrow{\Delta} 2Hg(l) + O_2(g)$, which reads – mercury(II) oxide is heated to give liquid mercury metal plus oxygen gas.

2) Metal hydrogen carbonates (bicarbonates) decompose to give a metal carbonate, water, and carbon dioxide.
Example: $2NaHCO_3(s) \xrightarrow{\Delta} Na_2CO_3(s) + CO_2(g) + H_2O(l)$, which reads – sodium bicarbonate is heated to give sodium carbonate, carbon dioxide gas, and water.

3) Metal carbonates decompose to give a metal oxide and carbon dioxide gas.

Example: $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$, which reads – calcium carbonate is heated to give calcium oxide plus carbon dioxide gas.

A **neutralization reaction** is the reaction between an **acid** and a **base**. Neutralization reactions are also called **acid-base** reactions. For this course, acids release hydrogen ions, H^+, while bases release hydroxide ions, OH^-. Example acids include HCl, HBr, H_2S, and H_2SO_4. We learned about them in Week 6. Example bases include NaOH, KOH, $Ca(OH)_2$, and $Ba(OH)_2$. Bases are made by reacting main group metal with the hydroxide ion. A neutralization reaction produces a **salt** and **water**.

Example acid-base reaction: $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
acid + base → salt + water

Acids are substances that release hydrogen ions, H^+, into water. Blue litmus paper turns red in the presence of hydrogen ions, so blue litmus paper turns red. Acids have a sour taste. Acids have a pH of less than 7 on the pH scale.

Bases are substances that releases hydroxide ions, OH^-, into water. Red litmus paper turns blue in the presence of hydroxide ions, so red litmus is used to test for bases. Bases have a slippery, soapy feel. Bases also have a bitter taste. Acids have a pH of greater than 7 on the pH scale.

Example neutralization reaction: $HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
acid + base → salt + water

The pH scale – most solutions have a pH between 0 and 14. The lower the pH, the more acidic is the solution. The higher the pH, the more basic is the solution. A neutral solution has a pH of exactly 7.

A **buffer** is a chemical system that resists changes in pH. Living systems depend on buffer systems to avoid extremes of pH. A buffer is a solution of a weak acid and one of its salts: Citric acid and sodium citrate make a buffer solution. When acid is added to the buffer, the citrate reacts with the acid to neutralize it. When base is added to the buffer, the citric acid reacts with the base to neutralize it.