Chapter 6 Lecture Notes: Chemical Reactions

Educational Goals

1. Define the term "**chemical reaction**."
2. Given the **reactants** and **products** in a chemical reaction, *write* and *balance* chemical equations.
3. Use **stoichiometric calculations** to determine the **theoretical yield** and **percent yield** of a reaction.
4. Identify **redox reactions** and determine which species is oxidized and which is reduced.
5. Understand and identify the four ***general types*** of reactions.
6. Identify combustion, hydrogenation of alkenes, hydrolysis of esters, hydration of alkenes, and dehydration of alcohol reactions.
7. Given the reactants for **combustion, hydrogenation of alkenes, hydrolysis of esters, hydration of alkenes,** or **dehydration of alcohol reactions**, be able to *predict* and *draw* structural formulas of the products.
8. Describe the difference in energy changes (ΔG) for **spontaneous** and nonspontaneous reactions, and list the factors that affect the **rate** of a chemical reaction.

In chapter 5 you learned about ***physical changes***. In chapter 6 you will learn about **chemical changes** that occur in processes called ***chemical reactions***. Chemical reactions occur in nature and in man-made events. A series of chemical reactions occurs in plants as they convert carbon dioxide and water molecules into the carbohydrate molecules that we eat and oxygen molecules that we breathe. Chemical reactions are used for propulsion in automobile engines and rocket thrusters. Some chemical reactions occur very quickly, as in explosions, and some reactions very slowly, as in the rusting of a nail. Chemistry is the study of how matter interacts with energy and/or other matter; *one way this happens is in chemical reactions*.

A chemical reaction is a process in which **chemical bond(s)** are *broken* **and/or** *new bonds are made*, such that one or more _______ ________________ are formed.

A Rocket Thruster

Example: One of the rocket engines in the space shuttle uses a chemical reaction in which oxygen gas and hydrogen gas are changed into gaseous H_2O molecules.

- To describe the chemical reaction, we say, "oxygen gas **reacts** with hydrogen gas to produce gaseous H_2O."

Scientists often write a ***chemical*** ___________ to describe a ***chemical reaction***.

- Chemical equations are similar to mathematical addition equations, except in chemical equations an arrow (→) is used *instead of* an equal (=) sign.
- For example, the **"*unbalanced*"** ***chemical equation*** for the reaction of oxygen and hydrogen to produce H_2O vapor is:

$$\mathbf{H_2 + O_2 \rightarrow H_2O}$$

When writing chemical equations, we often indicate the _________ of each reactant and product in parenthesis after its chemical formula.

- One of the following states is used: gas **(*g*)**, liquid **(*l*)**, solid **(*s*)**, or aqueous **(*aq*)**.
    - Aqueous **(*aq*)** indicates that the substance is ***dissolved in water***.
- For our rocket fuel example, we write:

$$\mathbf{H_2\ (\mathit{g}) + O_2\ (\mathit{g}) \rightarrow H_2O\ (\mathit{g})}$$

The substances on the ***left-hand side*** of the reaction arrow (→) are referred to as the ______________.

The substances on the ***right-hand side*** of the reaction arrow are referred to as the ______________.

In the rocket fuel chemical reaction, we began with two *reactants* (O_2 and H_2), and ended with one *product* (H_2O).

Antoine Lavoisier (1743-1794) and his wife, Marie-Anne Pierette Paulze (1758-1836)

Antoine Lavoisier and his wife, Marie-Anne Pierette Paulze, and Mikhail Lomonosov are credited for proposing and verifying the **law of conservation of mass**.

- This law states that ***matter is neither created nor destroyed in a chemical reaction, only the chemical bonding changes.***

The *law of conservation of mass* requires that the _________ ***number of atoms of each element*** appear ***on*** _________ _________ ***of the chemical equation***; when this is applied to a chemical equation, we say that the equation is "____________."

In our rocket fuel chemical equation, we need to have at least ***<u>two</u>*** H_2O molecules on the right-hand side to "balance" the two oxygen atoms on the left-hand side (in O_2).

We use a ________________ and re-write the equation as:

$$\mathbf{H_2\ (\mathit{g}) + O_2\ (\mathit{g}) \longrightarrow 2\ H_2O\ (\mathit{g})}$$

The **coefficients** indicate the **multiples** of each reactant and each product needed in order to have a *balanced equation*.

- If no coefficient appears for a substance, this implies that the coefficient is "___."

Because I added a coefficient of "**2**" to H_2O, the *oxygen atoms* are now *balanced*; there are *two oxygen* atoms ***on each side*** of the equation.

Is this equation completely balanced now?

$$H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$$

The equation is not yet balanced.

There are two hydrogen atoms on the reactant side (on the left) of the equation and the two H_2O molecules on the product side (right-hand side) contain a total of four hydrogen atoms (two hydrogens for <u>each</u> H_2O).

We need to use *another coefficient* to balance the equation.

- We need **two** H_2 molecules on the ***reactant side*** in order to balance the chemical equation:

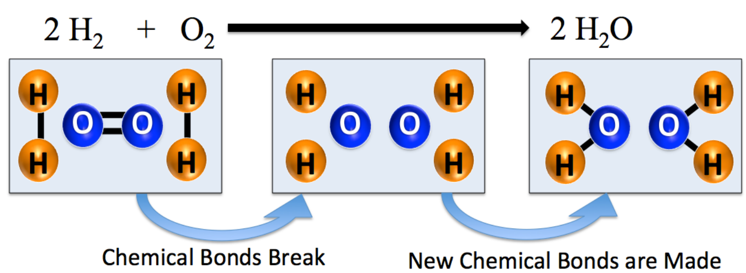
$$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$$

In this reaction, it takes ______ H_2 molecules to **react** with ______ O_2 molecule to **produce** ______ H_2O molecules.

I told you that a chemical reaction is a process in which ***chemical bond(s)*** are *broken* ***and/or*** new bonds are made such that one or more ***new substances*** are formed.

- Consider how this occurred in our reaction of rocket fuel.

In order to produce <u>two</u> H_2O molecules, ***two*** H-H bonds and the O-O double bond must break and ***four*** <u>new</u> O-H bonds must be made.



The exact order of bond breaking and making in chemical reactions is called the **reaction mechanism**. The details of the particular reaction mechanism for our reaction of rocket fuel and many other chemical reactions are not completely known and remain an active area of research in academia and industry.

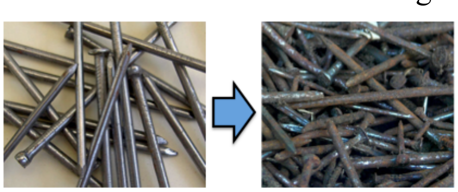
Observational Evidence of Chemical Reactions

When reactants are converted into new substances, several macro-scale observations may present evidence that a chemical reaction has occurred.

1) A ___________ ***change*** is evidence of a chemical reaction.

Many substances absorb visible light and therefore appear with a particular color. As a colored reactant is converted to product(s), the color of the reactant disappears. Chemical reactions ***always*** involve the formation of *one or more* ***new substances*** called ***product(s).*** If a product absorbs visible light, a new color will appear.

This is the basis of the color change seen in the chemical reaction that occurs when a nail "rusts."

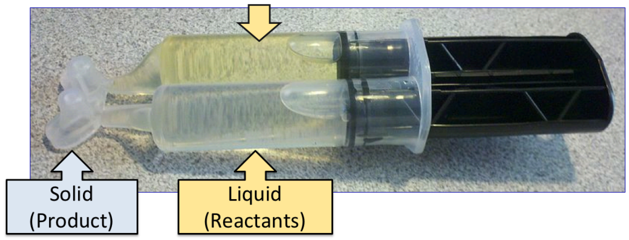


$4\ Fe\ (s) + 3\ O_2\ (g) \rightarrow 2\ Fe_2O_3\ (s)$

The reactants are iron metal and oxygen gas. Although iron metal does not absorb visible light, it does reflect it, and therefore has a "silver" appearance. The iron metal nail reacts with O_2 in the air to form Fe_2O_3 (rust). Fe_2O_3 absorbs all colors of visible light, a bit more strongly in the blue region, and therefore has a dark brown/red color.

2) The ***formation of a*** ______ _____________ is evidence of a chemical reaction.

- **Example:** Epoxy adhesives involve a chemical reaction of two ***liquid reactants*** that form a **new**, ***solid product*** substance.



- Another example of a reaction in which a ***new phase is formed*** is the reaction of aqueous acetic acid (white vinegar) with sodium bicarbonate (baking soda).

$$\underset{\textbf{acetic acid}}{C_3H_4O_2\,(aq)} + \underset{\textbf{sodium bicarbonate}}{Na_2HCO_3\,(s)} \longrightarrow \underset{\textbf{water}}{H_2O\,(l)} + \underset{\textbf{sodium acetate}}{Na_2C_3H_3O_2\,(aq)} + \underset{\textbf{carbon dioxide}}{CO_2\,(g)}$$

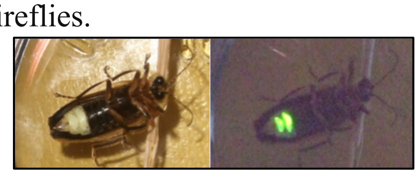
3) Observation of a ______________ ***change*** is evidence of a chemical reaction.

- **Example:** A "thermite" reaction.

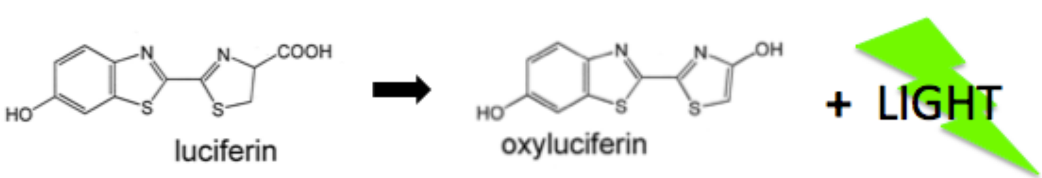
$$\underbrace{2\,Al\,(s) + Fe_2O_3\,(s)}_{\boxed{\text{Originally at room temperature.}}} \longrightarrow \underbrace{Al_2O_3\,(s) + 2\,Fe\,(l)}_{\boxed{\text{Products are very hot.}}}$$

4) Observation of the ***emission of*** ____________ is evidence of a chemical reaction.

- Examples of this are the thermite reaction that I just discussed and a chemical reaction that occurs in fireflies.



In a series of two reactions, a molecule called firefly luciferin is converted to oxyluciferin.



5) Observation of ***a new*** ______________ is evidence of a chemical reaction.

- An example of the detection of a new odor from a chemical reaction is "***rotting***."

6) The ***appearance of a*** _______________ is evidence of a chemical reaction.

- **Example:** The burning (combustion) of methane (CH_4), in natural gas.

$$CH_4\,(g) \quad + \quad 2\,O_2\,(g) \quad \longrightarrow \quad CO_2\,(g) \quad + \quad 2\,H_2O\,(g)$$

Understanding Check

Answer the following questions about the chemical equation shown below:

$$\mathbf{2\,H_2 + S_2 \rightarrow 2\,H_2S}$$

a. What are the reactants?

b. What is the product?

c. What is the number "2" in front of the H_2 (and H_2S) called?

d. Is the reaction balanced?

e. Why is there not a coefficient for S_2?

f. How many hydrogen **atoms** are needed to produce **two** H_2S molecules?

g. How many sulfur **atoms** are needed to produce **two** H_2S molecules?

h. How many hydrogen **molecules** are needed to produce **two** H_2S molecules?

i. How many sulfur **molecules** are needed to produce **two** H_2S molecules?

Balancing Chemical Equations

Matter is neither created nor destroyed in a chemical reaction.

- Therefore, the same __________ of atoms of each element appear on **both** ___________ of a ***balanced chemical equation***.

I introduced you to the ***law of conservation of mass*** and the need to balance chemical equations in the previous video.

- Some ***students*** are able to balance chemical equations "*in their heads*" by inspection of the unbalanced equation. This approach often involves trial and error placement of coefficients.
- Other students prefer a systematic, methodical approach.

In this video, I will provide you with a systematic method for balancing equations.

Whether you use the inspection (trial and error) approach *or* the systematic method, the end result is the *same balanced equation*.

There are ***three steps*** involved in the *systematic balancing method*:

Step 1: Make a table that lists the elements that are present and count all atoms on each side of the *unbalanced* equation.

- If **H_2 or O_2** is present, list these elements last.
- A polyatomic ion may be counted as one "element" ***if it appears on* both *sides of the equation.***

Step 2: Balance an element in the table by adding ***coefficient(s)*** to the equation (start with the first element on the list).

Step 3: Recount each atom and update the table, then repeat **Steps 2 and 3** for all elements as needed until the equation is balanced.

We will go right to an example problem where we will use these steps.

Take notes here – fill in the table and coefficients while watching the video.

Example: Balance the equation for the following reaction:

$$\_\_\_\_\_\_ \ H_2\,(g) + \ \_\_\_\_\_\_ \ O_2\,(g) \ \rightarrow \ \_\_\_\_\_\_\_ \ H_2O\,(g)$$

Amount on Reactant Side	Element	Amount on Product Side
	O	
	H	

Take notes here – fill in the table and coefficients while watching the video.

Example: Balance the following chemical equation:

$$\_\_\_\_\_\_ \ Mg\,(s) + \ \_\_\_\_\_\_ \ O_2\,(g) \ \rightarrow \ \_\_\_\_\_\_\_ \ MgO\,(s)$$

Amount on Reactant Side	Element	Amount on Product Side
	Mg	
	O	

Next, I want to show you and example problem that illustrates the short-cut that you can use when the same **polyatomic ion** appears on the reaction side **and** product side of the equation.

When I gave you the three steps involved in the systematic balancing method, I gave you a couple of instructions for performing "**Step 1**" that will save time and make balancing a bit simpler.

Step 1: Make a table that lists the elements that are present and count all atoms on each side of the *unbalanced* equation.

- If **H_2 or O_2** is present, list these elements last.
- A polyatomic ion may be counted as one "element" ***if it appears on <u>both</u> sides of the equation.***

Let's take a look at how this *second instruction* works.

Take notes here – fill in the table and coefficients while watching the video.

Example: Balance the following chemical equation.

$$\_\_\_\_\_\ Al\,(s) + \_\_\_\_\_\ FeSO_4\,(aq) \rightarrow \_\_\_\_\_\_\ Al_2(SO_4)_3\,(aq) + \_\_\_\_\_\ Fe\,(s)$$

Amount on Reactant Side	Element	Amount on Product Side
	(SO_4)	
	Al	
	Fe	

Step 1: Make a table that lists the elements that are present and count all atoms on each side of the *unbalanced* equation.

- If **H_2 or O_2** is present, list these elements last.

When **H_2** or **O_2** are present, we list hydrogen and oxygen last because **H_2** and **O_2** *consist of only one type of element*, and therefore when a coefficient is added to these molecules, it will not change the number of atoms of other elements that have already been balanced.

Take notes here – fill in the table and coefficients while watching the video.

Example: Balance the following chemical equation.

$$\_\_\_\_ \ C_3H_8\,(g) + \_\_\_\_\_ \ O_2\,(g) \rightarrow \_\_\_\_\_ \ CO_2\,(g) + \_\_\_\_\_ \ H_2O\,(g)$$

Amount on Reactant Side	Element	Amount on Product Side
	C	
	H	
	O	

Sometimes two coefficients must be __________________ applied in order to balance an element.

Take notes here – fill in the table and coefficients while watching the video.

Example: Balance the following chemical equation.

$$\_\_\_\_\_\_ \ Al\,(s) + \_\_\_\_\_ \ O_2\,(g) \rightarrow \_\_\_\_ \ Al_2O_3\,(s)$$

Amount on Reactant Side	Element	Amount on Product Side
	Al	
	O	

Common Errors to Avoid When Balancing Chemical Equations

When you are asked to balance an equation, *for example*:

$$\_\_\_\_\_\ N_2 + \_\_\_\_\_\ O_2 \rightarrow \_\_\_\_\_\ N_2O$$

Avoid making the following mistakes:

1. **Do not change** the ***formula*** of a reactant or product:

$$N_2 + O_2 \rightarrow N_2O_2$$

- By ***changing*** the product from N_2O to N_2O_2, you are not balancing the equation *for the given reaction*. The formation of N_2O_2 may or may not occur, however *it is not the reaction whose equation you were asked to balance*.

2) **Do not** add ***new*** reactants or products.

$$N_2 + O_2 \rightarrow N_2O + O$$

- By adding a new product (**O**), you are not balancing the equation for *the reaction that you were asked to balance*.

3) **Do not** use ***multiples of the coefficients*** when writing the balanced equation.

We ***would not*** write: $4\ N_2 + 2\ O_2 \rightarrow 4\ N_2O$

- Although the equation above is balanced, the convention for balancing equations is to use the ***lowest set*** of coefficients. To get this set, we divide each coefficient by the greatest common factor of all coefficients. The greatest common factor of the coefficients in our equation is **2**, so the correct balanced equation is written as:

$$2\ N_2 + O_2 \rightarrow 2\ N_2O$$

4) When chemical reactions are described in ***words***, you should identify the following ***diatomic molecules*** that are referred to by ***their element's names***:

$$H_2, N_2, O_2, F_2, Cl_2, Br_2, \text{ and } I_2$$

When uncombined with other elements in compounds, these elements exist as *diatomic molecules*.

- These elements are highlighted in the periodic table shown below.
- Except for H_2, these substances can be remembered by the upside-down "**L**" pattern seen in their periodic table positions.

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt									

When describing our example reaction, $2\ N_2 + O_2 \rightarrow 2\ N_2O$, ***in words***, we would say,

"***nitrogen*** reacts with ***oxygen*** to produce (or form) dinitrogen monoxide."

- Note that oxygen ***also*** exists as O_3, however O_3 is called "ozone."

Understanding Check

Balance the following chemical equations:

a. $Fe + O_2 \rightarrow Fe_2O_3$

b. $H_2 + Cl_2 \rightarrow HCl$

c. $Ag + H_2S \rightarrow Ag_2S + H_2$

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

e. $HgO \rightarrow Hg + O_2$

f. $Co + H_2O \rightarrow Co_2O_3 + H_2$

Understanding Check

Write ***balanced chemical equations*** for each of the following *equation descriptions*. You **do not** need to include the *states* of the reactants or products.

IMPORTANT NOTE: Before attempting to balance the equations, you must first convert the ***compound names*** into the correct ***chemical formulas***. If you begin to struggle with that, you may wish to go back to chapter 3 and re-work the naming problems.

a. Aluminum metal ***reacts with*** copper(II) chloride ***to produce*** aluminum chloride <u>and</u> copper metal.

b. Lead(II) nitrate reacts with sodium bromide to produce lead(II) bromide and sodium nitrate.

- HINT: Since you have the same polyatomic ion (nitrate) on both sides of the equation you can use the shortcut that was listed in Step #1 of our balancing method.

c. Zinc metal reacts with oxygen gas to produce zinc oxide.

- Oxygen is one of the *diatomic molecules* that are referred to by their *element's name*.

d. Aluminum sulfate reacts with barium iodide to produce aluminum iodide and barium sulfate.

e. At temperatures reached during baking, sodium bicarbonate (baking soda) decomposes (reacts) to produce sodium carbonate, carbon dioxide, and dihydrogen monoxide.

- Bicarbonate is a *polyatomic ion* (**not** the same as the carbonate polyatomic ion).

f. Sodium metal reacts with water to produce sodium hydroxide and hydrogen gas.

g. Lead(IV) sulfide reacts with oxygen gas to produce lead(IV) oxide and sulfur dioxide.

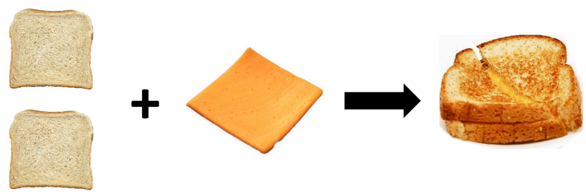
Stoichiometric Calculations for Chemical Reactions

Stoichiometry is the calculation of the ___________ of reactants and/or products in a chemical reaction.

The calculated values allow us to predict how much product(s) will be produced from a given amount of reactant(s), or conversely, how much reactant(s) will be needed in order to produce a desired amount of product.

Before we do ***stoichiometric calculations*** with chemicals, let's do a similar problem with something with which we are all familiar - ***food***!!!

I will use the formation of a grilled cheese sandwich as an *analogy* to a chemical reaction.



2 Slices of Bread + 1 Slice of Cheese ⟶ 1 Sandwich

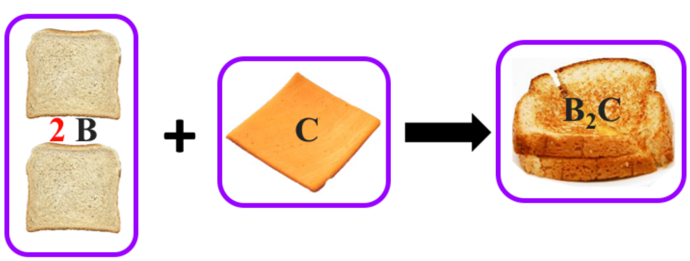
If this was an actual chemical reaction we would say, "two slices of bread ***react with*** one slice of cheese ***to produce*** one cheese sandwich."

Question #1*:* Suppose you want to make as many cheese sandwiches as possible for lunch. If you have 20 slices of bread, how many slices of cheese do you need? ______________

Question #2*:* If you have an *unlimited* supply of cheese slices, how may sandwiches can you make from the 20 slices of bread? ______________

When we do stoichiometry calculations for real reactions, the first thing that we need to have is a balanced chemical equation. For the sandwich analogy, the balanced equation is shown below.

$$\mathbf{2\,B + C \longrightarrow B_2C}$$



B represents a slice of bread, **C** represents a slice of cheese, and a sandwich "compound" is therefore a "$\mathbf{B_2C}$."

The ________________ between the amounts of reactants used and/or products formed can be found from the ________________ in the balanced chemical equations.

- It is for this reason that coefficients are sometimes referred to as "**stoichiometric coefficients**."

We treat **stoichiometry problems** like **unit conversation** problems.

- In ***Question #1***, you converted from ***units*** of "slices of bread" to ***units*** of "slices of cheese."

  Question #1*:* Suppose you want to make as many cheese sandwiches as possible for lunch. If you have 20 slices of bread, how many slices of cheese do you need? **10 slices of cheese**

$$20\ \cancel{\text{slices of bread}} \times \frac{1\ \text{slice of cheese}}{2\ \cancel{\text{slices of bread}}} = 10\ \text{slices of cheese}$$

- In ***Question #2***, you converted from ***units*** of "slices of bread" to ***units*** of "sandwiches."

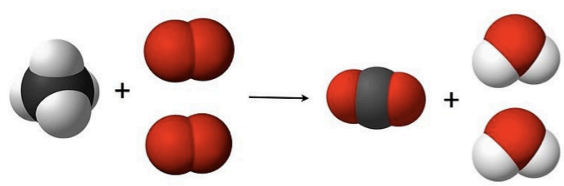
  Question #2*:* If you have an *unlimited* supply of cheese slices, how may sandwiches can you make from the 20 slices of bread? **10 sandwiches**

$$20\ \cancel{\text{slices of bread}} \times \frac{1\ \text{sandwich}}{2\ \cancel{\text{slices of bread}}} = 10\ \text{sandwiches}$$

I used the formation of a cheese sandwich analogy because you are quite familiar with making sandwiches; now we will move on to work on ***real chemical reactions***.

The coefficients in the balanced equation represent the ratios in which reactants are consumed and products are made in a chemical reaction.

- For example, consider the reaction for the combustion of methane.



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

- Use the balanced chemical equation to answer these questions:

  1) How many H_2O molecules are produced from 1 CH_4 molecule? ________
  2) How many O_2 molecules are needed to produce 2 H_2O molecules? ________
  3) How many CH_4 molecules are needed to react with 2 O_2 molecules? ________

In practice, scientists and technicians usually work with very large numbers of particles, not just a few molecules or ions.

It is more practical to consider the coefficients to represent the ratios in which __________ ***of substances*** are used and produced in a chemical reaction.

1) How many ***moles*** of H_2O are produced from **1 *mole*** of CH_4? **<u>2 moles</u>**
2) How many ***moles*** of O_2 are needed to produce **2 *moles*** of H_2O? **<u>2 moles</u>**
3) How many ***moles*** of CH_4 are needed to react with **2 *moles*** of O_2? **<u>1 mole</u>**

Your understanding of the coefficients in chemical equations will help you to construct ***conversion factors*** and use them in stoichiometric calculations.

Example: For the combustion of methane reaction, how many moles of H_2O can be produced from 2.84 moles of methane (CH_4)? Assume you have an unlimited supply of O_2.

- We will approach stoichiometry problems ***just as we did with*** _________ _______________ ***problems using our factor-label method***.
- In this problem, we are converting from units of "**moles of CH_4**" to units of "**moles of H_2O.**"

Implied Coefficient Implied Coefficient

$$\boxed{1}\,CH_4\,(g) + 2\,O_2\,(g) \rightarrow 1\,CO_2\,(g) + \boxed{2}\,H_2O\,(g)$$

$$\begin{array}{c|c|}
2.84\ \cancel{\text{moles CH}_4} & \boxed{2}\ \text{moles H}_2\text{O} \\ \hline
 & \boxed{1}\ \cancel{\text{mole CH}_4}
\end{array} = \mathbf{5.68}\ \text{moles H}_2\text{O}$$

In some stoichiometry problems, like the example problem just completed, you will be given a certain amount of one reactant and an *unlimited supply* of the other reactant(s), and then asked to calculate how much *product* could be made.

In other stoichiometry problems, you may need to calculate how much of one reactant *would be needed* to react with a given amount of ***another reactant*** as in the example that follows.

Example: For the combustion of methane reaction, how many moles of O_2 will be needed to react with 6.93 moles of CH_4?

- In this problem, we are converting from units of "**moles of CH_4**" to units of "**moles of O_2**."
- Get the relationship between **moles of CH_4** and **moles of O_2** ***from the*** **coefficients** ***in the chemical equation***.

Implied Coefficient Implied Coefficient

$$\boxed{1}\,CH_4\,(g) + \boxed{2}\,O_2\,(g) \rightarrow 1\,CO_2\,(g) + 2\,H_2O\,(g)$$

$$\begin{array}{c|c|}
6.93\ \cancel{\text{moles CH}_4} & \boxed{2}\ \text{moles O}_2 \\ \hline
 & \boxed{1}\ \cancel{\text{mole CH}_4}
\end{array} = \mathbf{13.9}\ \text{moles O}_2$$

Sometimes stoichiometry problems involve the calculation of how much of **one reactant** *would be needed* to ***produce*** a specified amount of ***product***.

Example: For the combustion of methane reaction, how many moles of O_2 will be needed to ***produce*** 1.74 moles of H_2O? Assume that there is an unlimited supply of CH_4.

- In this problem, we are converting from units of "**moles of H_2O**" to units of "**moles of O_2**."

$$CH_4\,(g) + \boxed{2}\,O_2\,(g) \rightarrow CO_2\,(g) + \boxed{2}\,H_2O\,(g)$$

$$\begin{array}{c|c|}
1.74\ \cancel{\text{moles H}_2\text{O}} & \boxed{2}\ \text{moles O}_2 \\ \hline
 & \boxed{2}\ \cancel{\text{moles H}_2\text{O}}
\end{array} = \mathbf{1.74}\ \text{moles O}_2$$

Understanding Check

In a combustion reaction, *propane* (C_3H_8) reacts with *oxygen* to produce CO_2 and H_2O.

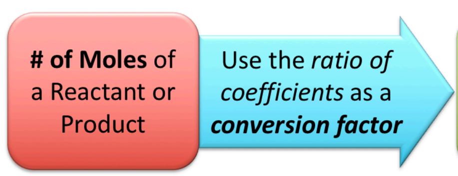
$$C_3H_8\,(g) + \mathbf{5}\,O_2\,(g) \rightarrow \mathbf{3}\,CO_2\,(g) + \mathbf{4}\,H_2O\,(g)$$

a) How many moles of H_2O can be produced from 1.32 moles of propane (C_3H_8)?
   - Assume there is an unlimited supply of O_2.

b) How many moles of O_2 will be needed to react with 12.6 moles of C_3H_8?

c) How many moles of O_2 will be needed to ***produce*** 0.843 moles of H_2O?
   - Assume there is an unlimited supply of C_3H_8.

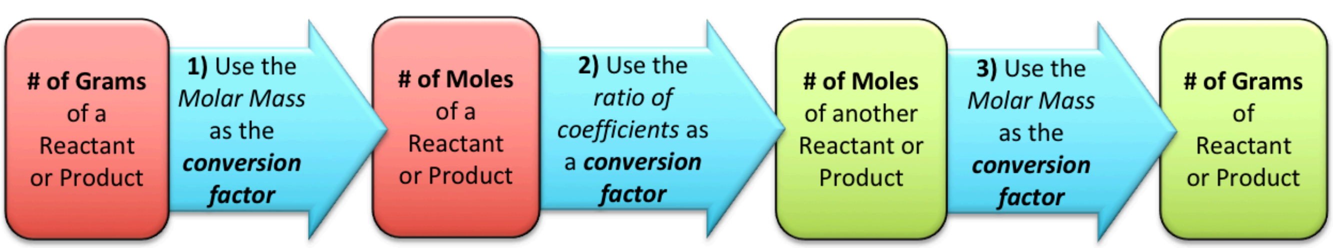
In the previous stoichiometry problems, we were given the number of moles of a particular product or reactant and asked to calculate the number of moles of another product or reactant involved in the reaction. We did so by using the ratios of the substances' coefficients as conversion factors. This process is shown in the schematic diagram below.



We cannot *directly* measure the number of moles on a balance in the lab.

In practice, we usually know the ***number of*** _________ of a reactant or product and wish to determine the ***number of grams*** of another reactant needed ***or*** the ***number of grams*** of a product in a reaction.

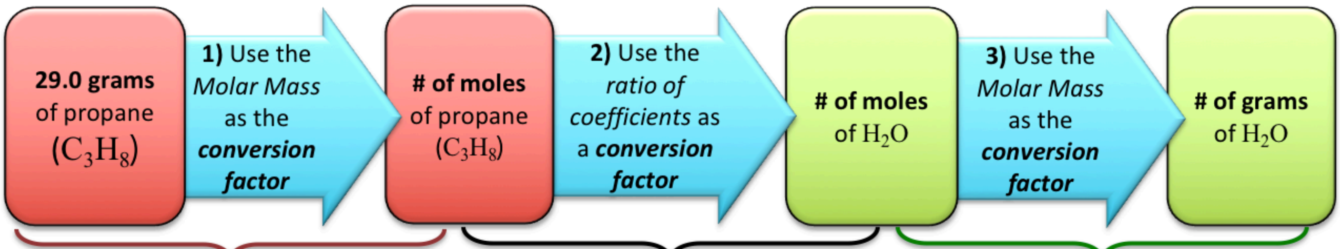
Because the coefficients that we use as conversion factors in stoichiometric calculations are the ***ratios of moles***, we must **first convert all masses (grams) to moles,** do the stoichiometric calculations, and then convert the calculated number of moles for the substance of interest into grams. This ***three-step process*** is shown in the schematic diagram below.



Example: For the combustion of *propane*: $C_3H_8\,(g) + \mathbf{5}\,O_2(g) \rightarrow \mathbf{3}\,CO_2\,(g) + \mathbf{4}\,H_2O\,(g)$

How many **grams** of H_2O can be produced from 29.0 **grams** of propane (C_3H_8)?

- Assume you have an unlimited supply of O_2.



Step # 1:
Convert from grams of C_3H_8 to moles of C_3H_8.
- Use the ***molar mass*** of C_3H_8:

$$\frac{29.0\ \cancel{\text{grams } C_3H_8}}{} \left|\, \frac{\mathbf{1}\ \text{mole } C_3H_8}{44.11\ \cancel{\text{grams } C_3H_8}} \right| = \mathbf{0.657\ moles\ C_3H_8}$$

Step #2:
Convert from moles of C_3H_8 to moles of H_2O.
- Use the *ratio of coefficients* as a ***conversion factor***

$$\frac{\mathbf{0.657}\ \cancel{\textbf{moles } C_3H_8}}{} \left|\, \frac{\mathbf{4}\ \text{moles } H_2O}{\mathbf{1}\ \cancel{\text{mole } C_3H_8}} \right| = \mathbf{2.63\ moles\ H_2O}$$

Step #3:
Convert from moles of H_2O to grams of H_2O.
- Use the ***molar mass*** of H_2O:

$$\frac{\mathbf{2.63}\ \cancel{\textbf{moles } H_2O}}{} \left|\, \frac{18.02\ \text{grams } H_2O}{1\ \cancel{\text{mole } H_2O}} \right| = \mathbf{47.4\ grams\ H_2O}$$

Alternative Solution: You may feel comfortable enough to use the short cut where all three of the conversions above are combined into one equation:

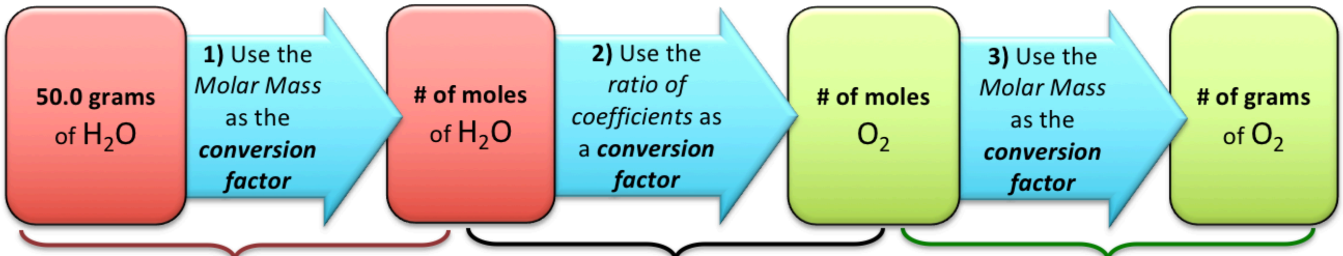
Step # 1 **Step # 2** **Step # 3**

$$\frac{29.0\ \cancel{\text{grams } C_3H_8}}{} \left|\, \frac{\mathbf{1}\ \cancel{\text{mole } C_3H_8}}{44.11\ \cancel{\text{grams } C_3H_8}} \right|\, \frac{\mathbf{4}\ \cancel{\text{moles } H_2O}}{\mathbf{1}\ \cancel{\text{mole } C_3H_8}} \left|\, \frac{18.02\ \text{grams } H_2O}{1\ \cancel{\text{mole } H_2O}} \right| = \mathbf{47.4\ grams\ H_2O}$$

Example: For the combustion of *propane*: $C_3H_8\,(g) + \mathbf{5}\,O_2(g) \rightarrow \mathbf{3}\,CO_2\,(g) + \mathbf{4}\,H_2O\,(g)$

How many ***grams*** of O_2 will be needed to ***produce*** 50.0 grams of H_2O?

- Assume that you have an unlimited supply of C_3H_8.



Step # 1:
Convert from grams of H_2O to moles of H_2O.
- Use the ***molar mass*** of H_2O:

$$50.0\ \cancel{\text{grams } H_2O} \times \frac{\mathbf{1}\ \text{mole } H_2O}{18.02\ \cancel{\text{grams } H_2O}} = \mathbf{2.77\ moles\ H_2O}$$

Step #2:
Convert from moles of H_2O to moles of O_2.
- Use the *ratio of coefficients* as a ***conversion factor***

$$\mathbf{2.77}\ \cancel{\textbf{moles } \mathbf{H_2O}} \times \frac{\mathbf{5}\ \text{moles } O_2}{\mathbf{4}\ \cancel{\text{moles } H_2O}} = \mathbf{3.46\ moles\ O_2}$$

Step #3:
Convert from moles of O_2 to grams of O_2.
- Use the ***molar mass*** of O_2:

$$\mathbf{3.46}\ \cancel{\textbf{moles } \mathbf{O_2}} \times \frac{32.00\ \text{grams } O_2}{1\ \cancel{\text{mole } O_2}} = \mathbf{111\ grams\ O_2}$$

Alternative Solution: You may feel comfortable enough to use the short cut where all three of the conversions above are combined into one equation:

Step # 1 **Step # 2** **Step # 3**

$$50.0\ \cancel{\text{grams } H_2O} \times \frac{\mathbf{1}\ \cancel{\text{mole } H_2O}}{18.02\ \cancel{\text{grams } H_2O}} \times \frac{\mathbf{5}\ \cancel{\text{moles } O_2}}{\mathbf{4}\ \cancel{\text{moles } H_2O}} \times \frac{32.00\ \text{grams } O_2}{1\ \cancel{\text{mole } O_2}} = \mathbf{111\ grams\ O_2}$$

Understanding Check

For the combustion of *propane*: $C_3H_8\,(g) + \mathbf{5}\,O_2(g) \rightarrow \mathbf{3}\,CO_2\,(g) + \mathbf{4}\,H_2O\,(g)$

How many ***grams*** of O_2 will be needed to ***react with*** 70.0 ***grams*** of C_3H_8?

Energy Changes in Chemical Reactions

All chemical reactions involve changes in __________.

Some reactions ___________ ***energy*** as heat, light, electricity, and/or mechanical energy (work).

The energy that is released in a chemical reaction comes from ____________ ***energy*** contained in the _____________.

Examples of reactions that produce heat and light are combustion reactions (burning).

- We call reactions that release energy, in the form of heat, _______________ **reactions**.

Combustion reactions can also produce *mechanical energy; another way to state this is, "combustion reactions can do work."*

- For example, when combustion occurs in the cylinder/piston system of an internal combustion engine, a sudden increase in the number of moles of gas present (produced in the reaction) causes a large increase in pressure within the cylinder, which then moves the piston. The piston is coupled to a shaft that rotates and ultimately rotates the wheels. Examples of reactions that release electrical energy are the reactions that occur in various types of batteries.

Some reactions must continuously ___________ ***energy*** in order to occur.

- **Example:** The *formation* of hydrogen and oxygen gas (at room temperature) from water.

$$2\ H_2O\ (l) \longrightarrow 2\ H_2\,(g)\ +\ O_2\,(g)$$

- Note that this is the ***reverse*** of the rocket fuel reaction that I discussed earlier. The conversion of oxygen and hydrogen to H_2O releases energy; so it makes sense that the ***reverse reaction***, converting H_2O to hydrogen and oxygen would require energy.

- The conversion of water to hydrogen and oxygen gas can be done by adding electrical energy to water in a process called **electrolysis**.

In a battery that is submerged in water, H_2O (l) is converted to O_2 (g) at one terminal and H_2 (g) at the other terminal.

Spontaneity of Chemical Reactions

Recall the important law that is central to understanding nature: **matter tends to exist in the lowest possible energy state**.

We applied this law to understand why light is emitted from atoms and why most substances exist as molecules or ions instead of as isolated atoms.

This is a *universal law,* therefore it also applies to chemical reactions.

Chemical reactions can occur when the *total energy of the products* is less than the *total energy of the reactants*.

When a chemical reaction can continue to occur *without an external input of energy*, we say the reaction is ____________________.

Let's examine the concepts of reaction energy and spontaneity by examining a reaction that we are now familiar with - the combustion of propane gas.

$$C_3H_8\,(g)\ +\ 5\ O_2\,(g)\ \longrightarrow\ 3\ CO_2\,(g)\ +\ 4\ H_2O\ (g)$$

Many of us have used propane gas in outdoor lanterns and barbecues. We know that once we use a spark or match to start the reaction, that the combustion reaction continues to occur *without an external input of energy*. It is therefore a ***spontaneous reaction***.

Let's think about the change in energy (Δ**E**) in the combustion of propane.

The **total energy** (**E**) is the sum of ***kinetic energy*** ($\mathbf{E_k}$) and ***potential energy*** ($\mathbf{E_p}$).

$$\mathbf{E = E_k + E_p}$$

A particle's ***kinetic energy*** comes from its ***motion***.

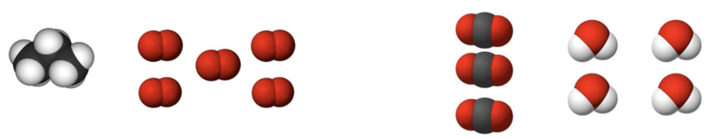
Potential energy is stored in a substance's ____________ ***bonds*** (and noncovalent interactions).

- _____________ chemical bonds (or noncovalent interactions) ___________ energy.
- _____________ new chemical bonds (or noncovalent interactions) ___________ energy.

For the combustion of propane reaction, in the ***reactants***, ***potential energy*** is stored in C-C bonds, C-H bonds, and the O-O bonds.

During the reaction, chemical bonds in the reactants are broken and new chemical bonds (C-O and O-H) are formed in the products.

$$C_3H_8\,(g)\ +\ 5\ O_2\,(g)\ \longrightarrow\ 3\ CO_2\,(g)\ +\ 4\ H_2O\ (g)$$



In the ***products***, ***potential energy*** is stored in C-O bonds, and O-O bonds.

We all know, from our familiarity with the combustion of propane gas, that energy is ***released*** in this reaction.

- The ***amount* of energy released** in the reaction is equal to the ____________ in ***potential energy*** ($\Delta\mathbf{E_p}$) between the products ($\mathbf{E_{p\ products}}$) and reactants ($\mathbf{E_{p\ reactants}}$):

Energy Released ($\Delta\mathbf{E_p}$) = ($\mathbf{E_{p\ products}}$) - ($\mathbf{E_{p\ reactants}}$)

This ***released energy*** does not just vanish into a hole in the universe! Energy is never created or destroyed, it only changes its form.

The energy released in the reaction is converted to ***kinetic energy in the products***, and then is often transferred to surrounding matter such as air (or the items cooking on the grill).

In the case of similar reactions that happen in *internal combustion engines*, some of the energy that is released goes into mechanical work (moving the car), and the remainder goes into warming of the engine, and then finally into warming of the air that contacts the car's engine and radiator. The ratio of energy that goes into *moving the car* to the total energy released by the reaction is called the **engine efficiency**. In the case of reactions that occur in batteries, most of the energy released in a reaction goes into kinetic energy of electrons moving through an electrical circuit; the remainder of the energy goes into warming the battery and surroundings.

The field of study called **thermodynamics** often deals with calculating the energy changes in chemical reactions. These energetic calculations are useful for predicting how chemical systems will behave under various conditions because matter tends to exist in the lowest possible energy state. For example, the combustion of propane reaction occurs ***because the bonding arrangement of atoms in the products is at a lower energy than that of the reactants.***

Energy calculations are only useful when they involve properties (variables) that can be directly measured (such as temperature, moles of each substance, volume, and pressure).

When using the temperature and pressure as variables in energy calculations, we refer to the calculated energy as "_________ _____________" abbreviated with the symbol "**G.**"

- The details of the difference between ***total energy*** (**E**) and ***free energy*** (**G**) are beyond the scope of this course.
- For *our purposes*, you can consider the ***free energy*** (**G**) and ***total energy*** (**E**) to be _______________.
- **Matter tends to exist in the lowest possible *free energy* state, therefore chemical reactions will occur when the total *free energy* of the products is less than the total *free energy* of the reactants**.

Let's consider the *combustion of propane* in terms of ***free energy*** **(G)**.

$$C_3H_8\,(g) \;+\; 5\,O_2\,(g) \;\longrightarrow\; 3\,CO_2\,(g) \;+\; 4\,H_2O\,(g)$$

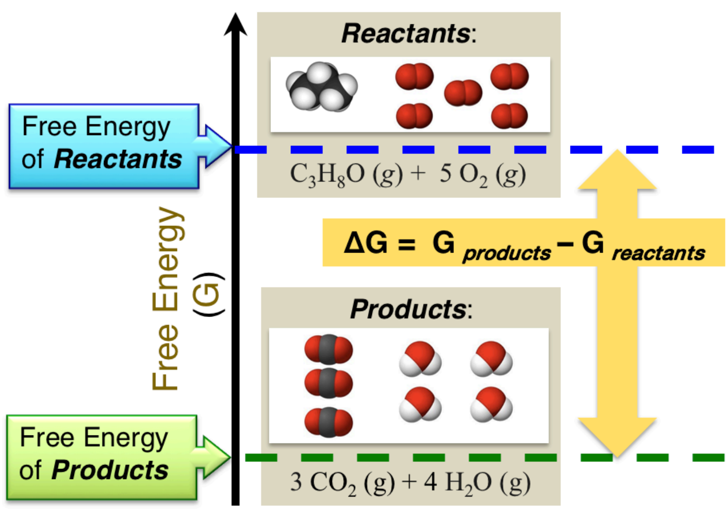
The figure shown below illustrates the difference in ***free energy*** between the products and the reactants for this reaction.

This figure is an *energy level diagram*, similar to the energy level diagrams that you saw in chapter 3, with higher energy toward the top.

The amount of free energy present in the reactants *alone* (**G***reactants*) is indicated by the upper dashed line (blue in the video).

The amount of free energy contained in products *alone* (**G***products*) is indicated by the lower dashed line (green in the video).

The ___________ ***in free energy*** (ΔG) for reaction is equal to the difference in free energy between the products (**G***products*) and the reactants (**G**reactants):



$$\Delta G = (G_{\text{products}}) - (G_{\text{reactants}})$$

Note the use of our convention of defining **change (Δ)** as the final state (products only) ***minus*** the initial state (reactants only).

The free energy of the products for the combustion of propane is **less than** the free energy of the reactants as indicated by their positions in the energy level diagram.

When the *free energy* of the ***products*** of a reaction is ***less than*** the free energy of the ***reactants***, we say that the reaction is **exergonic**.

Chemical reactions will occur spontaneously when the *free energy* of the product(s) is less than the *free energy* of the reactant(s).

$$\Delta G = (G_{\text{products}}) - (G_{\text{reactants}})$$

When the ***free energy*** of the *products* of a reaction is ***less than*** the free energy of the *reactants*, the ***change in free energy*** (Δ**G**) will have a ______________ ***value***.

Summary of Spontaneity of Reactions

When a chemical reaction can continue to occur *without an external input of energy*, we say the reaction is ***spontaneous***.

The following statements are true for ***spontaneous reactions***:

- The ***free energy*** (Δ**G**) the ***products*** of the reaction is ***less than*** the ***free energy*** of the ***reactants***.
- The sign of Δ**G** is ***negative***.

Understanding Check

Determine if the following reactions are **spontaneous** or **non-spontaneous**.

1. $2\ H_2\ (g) + O_2\ (g) \longrightarrow 2\ H_2O\ (g)$, Δ**G** = -54,640 calories (per mole of product formed)

2. $2\ H_2O\ (l) \longrightarrow 2\ H_2\ (g) + O_2\ (g)$, Δ**G** = 56,670 calories (per mole of products formed)

3. The rusting of an iron nail: $4\ Fe + 3\ O_2 \longrightarrow Fe_2O_3$
   (HINT: Have you ever observed a rusty nail?)

Rates of Chemical Reactions

Knowing if the value of Δ**G** is negative or positive allows us to predict whether or not a reaction is spontaneous, however it does not give us any information about ***how quickly a reaction happens***.

Some reactions happen ***very quickly***, for example, the explosions of fireworks.

Some reactions happen so ***slowly*** that you cannot tell that the reaction is occurring even when viewed for a few hours, for example the rusting of a nail.

To understand the factors that determine and influence how quickly chemical reactions happen, we must consider the energy during ***the process*** of converting reactants to products. We will do so using a graph of the free energy ***as the reaction progresses***.

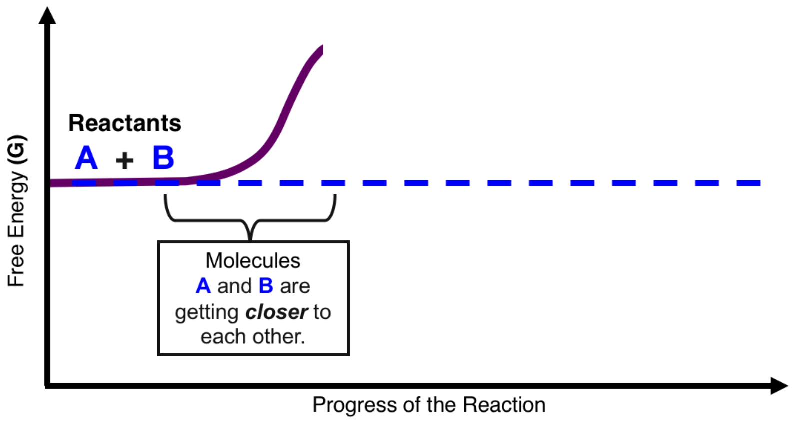
Let's consider a generalized chemical reaction where **molecule A** *reacts* with **molecule B** to *produce* **molecule C** and **molecule D**:

$$\mathbf{A + B \longrightarrow C + D}$$

A + B → C + D

We plot the free energy (G) on the vertical axis. The horizontal axis indicates the *progress of the reaction* (process of converting reactants to products). The **curve** (purple in the video) represents the free energy during the reaction process.

We begin on the ***left*** where we have the *reactants only*.



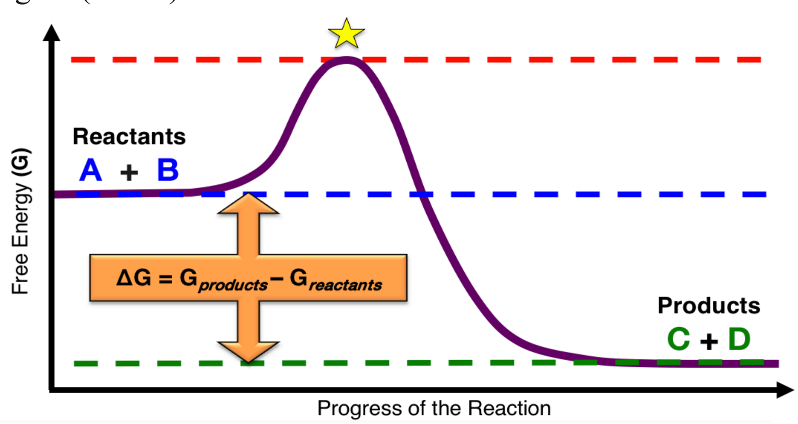
The free energy of the reactants ($G_{reactants}$) are indicated by the horizontal dashed line (blue in the video).

As the reaction progresses, we move from *left to right* on the horizontal axis. Originally molecule **A** and molecule **B** are fairly far apart from each other. As the reaction progresses, **A** and **B** must be on a collision course with each other. In order for molecule **A** to react with molecule **B**, they must eventually collide.

As **A** and **B** approach each other, ***the free energy increases*** because of factors such as electrostatic repulsion between the electrons on the reactants, changes in molecular geometry, and disruption of noncovalent interactions.

- This increase in energy can be seen in the free energy curve (purple in video) as the reaction progresses.

At some point, the free energy reaches a maximum value (peak) as indicated by the yellow star in the figure (below).



At this point, the matter no longer exists as individual reactant molecules (**A** and **B**), nor has it been converted to product molecules (**C** and **D**).

The matter exists in what we call a **transition state**, where the bonds in the reactants have not all been completely broken and/or the new bonds in the products have not been completely formed.

The ***free energy*** of the ***transition state*** is indicated by the top dashed line (red in the video).

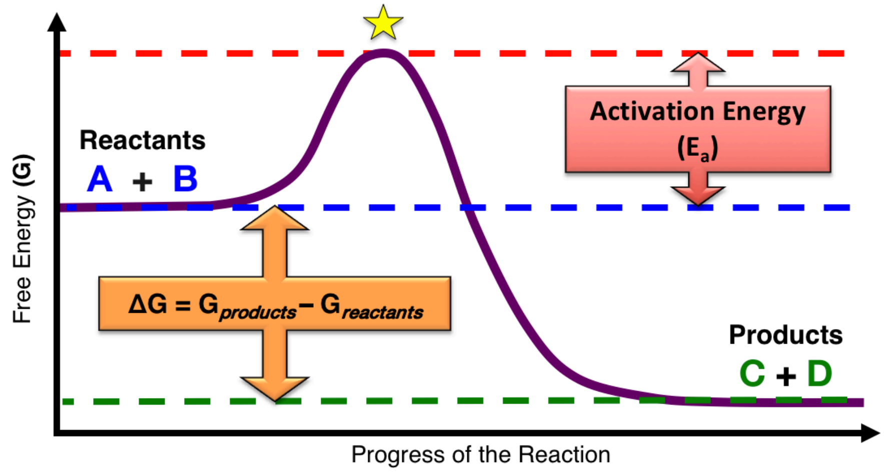
As the reaction progresses past the transition state, new bonds are formed in the product molecules **C** and **D**, and bond angles and bond distances relax to their low energy geometries, and the product molecules move apart from each other.

All of these processes result in ***lowering the free energy*** as can be seen in the figure as the reaction progresses from the transition state to the products.

- The free energy of the products is indicated by the bottom dashed line (green in the video).

In this example, the free energy of the products is less than the free energy of the reactants, therefore **ΔG** is ***negative*** and the reaction is ***spontaneous***.

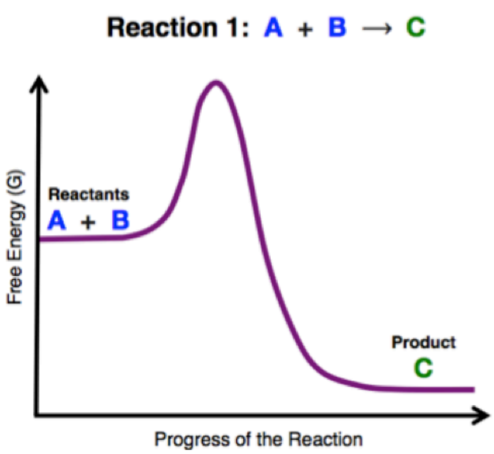
The amount of free energy needed to progress ***from reactants to the transition state*** is called the ______________ ____________ (**E_a**).

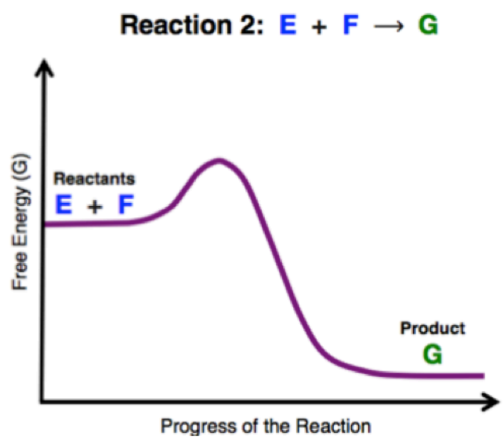


The _______ **of chemical reactions** (how quickly the reactions happen) depend on the ***activation energy***.

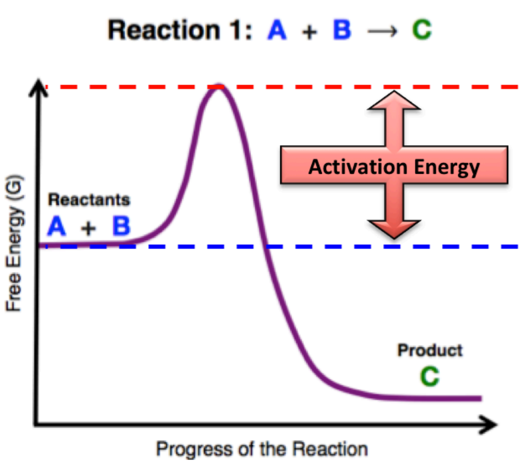
The _________ the activation energy, the _________ the reaction rate.

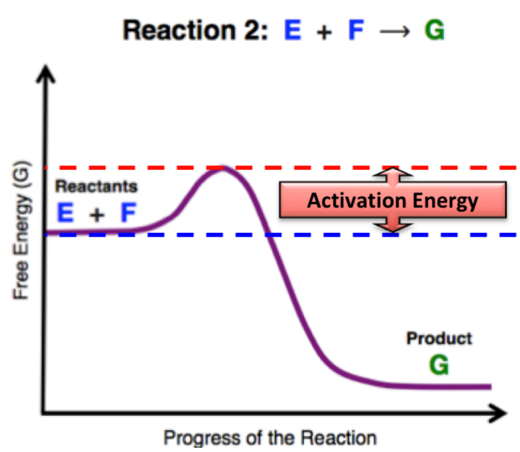
Example: The energy level diagrams for two *spontaneous* reactions are shown below. Which reaction has a ***faster rate***?





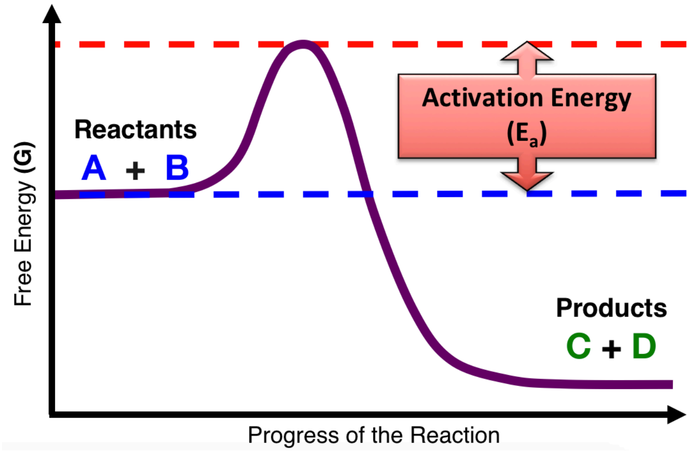
Solution: Compare the ***activation energies***. The ***lower*** the activation energy, the ***faster*** the reaction rate.





Because **Reaction 2** has a ***lower*** activation energy, it has the ***faster reaction rate***.

The Temperature Dependence of Reaction Rates



The reaction energy level diagram shows an increase in free energy as the reaction progresses from reactants to the maximum energy peak at the *transition state*.

You just learned that this increase in energy is referred to as the ***activation energy.***

If the temperature of the reactants is increased, the reactants can more readily overcome the *activation energy* and therefore more readily be converted to products.

For this reason, the reaction rate depends on the ______________.

- ***As the temperature increases, the reaction rate increases.***
    - In general, for every 10° C increase in temperature, the reaction rate doubles.
    - Conversely, for every 10 °C decrease in temperature, the reaction rate decreases by a factor of one-half.
        - This is one of the reasons that we keep food refrigerated; the lower temperature slows the reactions that are involved in the decomposition of food and bacterial growth.

Catalysis

Another way to change the rate of a chemical reaction is to use a ________________.

A ***catalyst*** can be any *substance that increases the rate of a chemical reaction.*

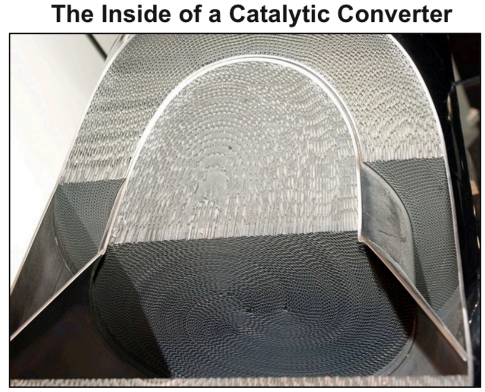
Unlike reactants, catalysts are ____ __________ in a reaction.

Living organisms produce catalysts consisting of large molecules, usually proteins, that are called ___________.

- Humans have thousands of chemical reactions that must occur in order to sustain life. Many of these reactions would happen too slowly to be useful if not catalyzed by enzymes. For example, an enzyme called amylase, present in our saliva, catalyzes the digestion reaction of starch.

Industrial processes often use ______________ of substances such as metals to catalyze reactions.

The Inside of a Catalytic Converter

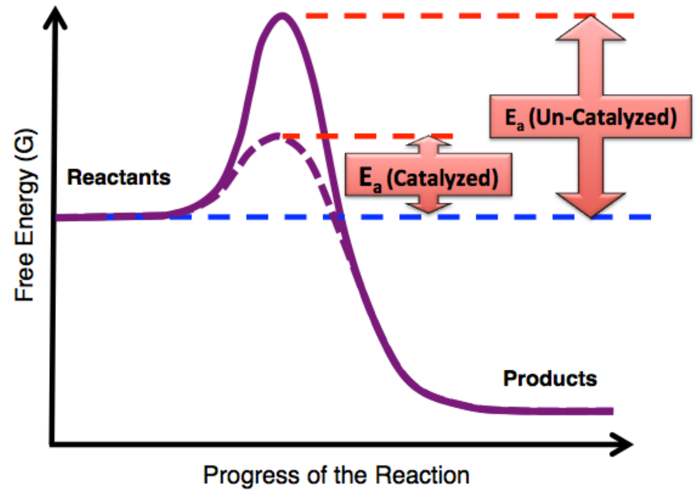


- For example, ***catalytic converters*** use platinum or rhodium surfaces as catalysts to remove poisonous by-products (carbon monoxide, nitrogen oxides, and unreacted hydrocarbons) produced in the ***incomplete*** combustion of fossil fuels.

Catalysts increase the rates of reactions by ______________ the _______________ ***energy***.

In the catalyzed reaction, the reactants require less energy to overcome the ***activation energy*** and are therefore converted to products at a ***faster*** rate.

You will learn more details of how catalysts lower the activation energy when I discuss enzymes in chapter 13.



Beach Balls in a Lake Analogy for Rates of Reactions

I want to end this video by giving you an analogous model to help you understand how temperature and catalysts effect the rate of chemical reactions. The process that I will use to model chemical reactions is beach

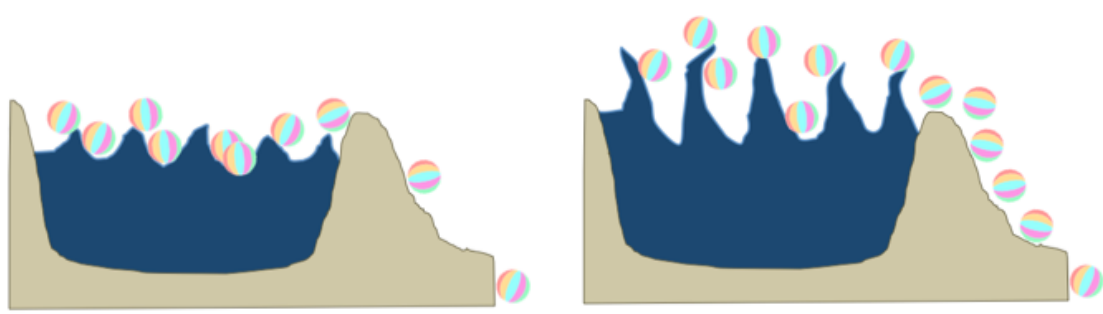
Beach balls (in the lake) $\longrightarrow$ Beach-balls (out of the lake)

The Temperature Dependence of the Reaction Rate

In this model, the *height of the lake's shoreline* represents the **activation energy**.

Increasing the size of the waves in the lake is analogous to raising the temperature of a chemical reaction.

Larger waves cause the beach-balls to leave the lake ***at a faster rate*** as shown in the illustration below:

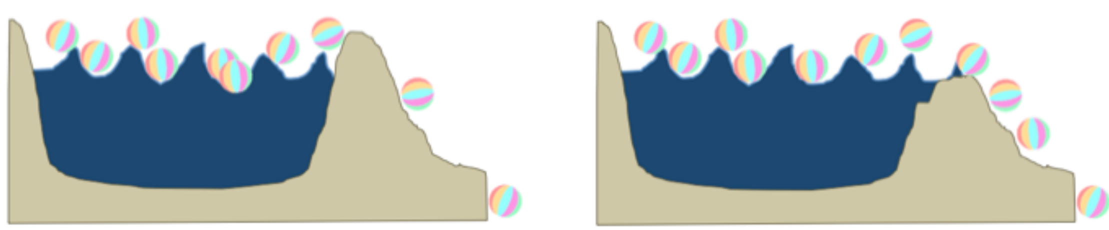


Lower Temperature Higher Temperature

The Effect of Catalysts on the Reaction Rate

Again, the *height of the lake's shoreline* represents the **activation energy**.

The *presence of a catalyst* is analogous to ***lowering the shoreline***, and results in the beach-balls leaving the lake *at* ***faster rate*** as shown in the illustration below:



Un-Catalyzed Reaction Catalyzed Reaction

Summary of Rates of Chemical Reactions

The **reaction rate** is a measure of *how **quickly** a reaction occurs*.

The **rates** of chemical reactions depend on the ***activation energy*** and the ***temperature***.

The ***lower*** the ***activation energy***, the ***faster*** the reaction rate.

As the ***temperature increases***, the reaction rate ***increases***.

- In general, the reaction rate **doubles** with every 10 ^{o}C increase in temperature.

Catalysts increase the rate of a reaction by ***decreasing*** the **activation energy (E_a)**.

General Types of Chemical Reactions

Many reactions can be categorized into one of ***four general types*** based solely on *changes in the bonding pattern* (not the identity of the reactants or products).

The ***educational goal*** for this section is that, given one of these four types of reactions, you will be able to name the category in which it belongs.

1) SYNTHESIS REACTIONS

A ______________ ***reaction*** is one in which a single compound is formed from two or more substances. The general form of a synthesis reaction is:

$$\mathbf{A + B \rightarrow AB}$$

Where **A** represents an element or compound, **B** represents another element or compound, and **AB** is the compound formed from **A** and **B**. An example of a synthesis reaction is the reaction that occurs between sodium metal and oxygen gas (O_2):

$$\mathbf{4\,Na}\,(s) + \mathbf{O_2}\,(g) \rightarrow 2\,\mathbf{Na_2O}\,(s)$$

2) DECOMPOSITION REACTIONS

A ____________________ ***reaction*** is a reaction in which a single reactant breaks down into two or more substances. The general form of a decomposition reaction is:

$$\mathbf{AB \rightarrow A + B}$$

- It is simply the reverse of a synthesis reaction.

An example of a ***decomposition reaction*** is the thermal (heat induced) decomposition of mercury(II) oxide:

$$2\,\mathbf{HgO}\,(s) \rightarrow 2\,\mathbf{Hg}\,(l) + \mathbf{O_2}\,(g)$$

Note that the key to identifying a decomposition reaction is that one reactant species is converted to two or more product species. In our example, we start the reaction with just one reactant present, HgO (s); after the reaction occurs, there are two different substances, Hg (l) and O_2 (g).

3) SINGLE-REPLACEMENT REACTIONS

In a ___________________________ ***reaction***, an element *replaces* another element from a compound.

The general form of a single-replacement reaction, where **A** replaces **B**, is:

$$\mathbf{A + BX \rightarrow AX + B}$$

A and **B** represent different *elements*, **BX** represents a compound made from **B** and **X**, and **AX** is the compound made of **A** and **X**. *Before reacting*, element **A** is alone and element **B** is in compound **BX.** *After the reaction,* element **B** is alone and element **A** is in compound **AX.**

An example of a single-replacement reaction is:

$$\mathbf{Cu}(s) + \mathbf{AgNO_3}\,(aq) \rightarrow \mathbf{Cu(NO_3)_2}\,(aq) + \mathbf{Ag}\,(s)$$

In this reaction, copper metal (Cu (*s*)) *replaces* the silver ion (Ag^+) in silver nitrate.

4) DOUBLE-REPLACEMENT REACTIONS

In a ___________________________ ***reaction***, two substances "***switch partners***." The general form of a single replacement reaction, where **AX** and **BY** *switch partners*, is:

$$\mathbf{AX + BY \rightarrow AY + BX}$$

Double-replacement reactions occur in *aqueous* solutions. You will learn more about double-replacement reactions when I discuss solutions in chapter 7. An example of a double-replacement reaction is the reaction of sodium chloride and silver nitrate:

$$\mathbf{NaCl}\,(aq) + \mathbf{AgNO_3}\,(aq) \rightarrow \mathbf{NaNO_3}\,(aq) + \mathbf{AgCl}\,(s)$$

Understanding Check

Categorize each of the following reactions as either: synthesis, decomposition, single-replacement, or double-replacement.

a. $2\,H_2O\,(l) \rightarrow 2\,H_2\,(g) + O_2\,(g)$

b. $KBr\,(aq) + AgNO_3\,(aq) \rightarrow KNO_3\,(aq) + AgBr\,(s)$

c. $2\,Mg\,(s) + O_2\,(g) \rightarrow 2\,MgO\,(s)$

d. $Mg\,(s) + 2\,HCl\,(aq) \rightarrow MgCl_2\,(aq) + H_2\,(g)$

Redox Reactions

The term "***redox***" is an abbreviated combination (*portmanteau*) of the words "______________" and "______________."

In a ***redox*** chemical reaction, an ***oxidation*** and a ***reduction*** occur ___________________.

Many of the reactions that I have used as examples in previous sections are redox reactions. Redox reactions often occur in biological systems. For example, the series of chemical reactions in which we *metabolize food* and the series of chemical reactions called *photosynthesis* both contain many redox reactions.

Oxidation is the _________ ***of*** _________________***(s)*** by an atom, ion, or molecule.

Reduction is the __________ ***of electron(s)*** by an atom, ion, or molecule.

In a redox reaction, **electrons are** ___________________ **from one atom, ion, or molecule *to another* atom, ion, or molecule.**

- The electron(s) that are "lost" by the *oxidized species* are "gained" by the *reduced species*.

A useful mnemonic to differentiate oxidation and reduction is the term "**OILRIG**"

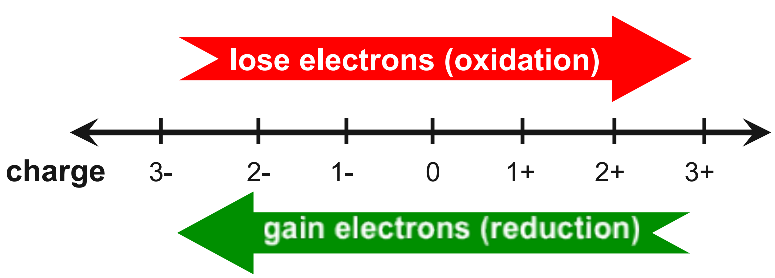
(**O**xidation **i**s the **L**oss of electrons; **R**eduction **i**s the **G**ain of electrons).

Redox Reactions of *Inorganic* Compounds

It is possible to identify redox reactions for ________________ ***compounds*** by inspecting the chemical equation and determining if electrons are transferred from one species to another.

- One atom or ion in a reactant will **lose electron(s)** and therefore its _____________ will **increase** (in the positive direction) by one charge unit for every electron that is lost.
  - The electron(s) that are lost are transferred to another reactant.

- The **charge** of the atom or ion in the reactant that **gains** the electron(s) will **decrease** (toward negative values) by one charge unit for every electron that is gained.



Example: Let's consider the reaction of sodium metal and oxygen:

$$4\ Na\ (s)\ +\ O_2\ (g)\ \rightarrow\ 2\ Na_2O\ (s)$$

To determine if electrons are transferred from one species to another, you must understand that ***the*** **_________ _________ *of any pure* ____________ *or* ______________ *is always* __________*!***

- This fact will help you to determine the charge of each atom or ion in the reactants and products, and then to know if the charge of the species ***changed*** during the reaction.
- This change in the charge of an atom or ion is an indication that a redox reaction occurred.

FILL IN THE TABLE BELOW AS I DO SO IN THE VIDEO:

Oxidation *or* Reduction	Reactant	Charge in Reactant	Element	Product	Charge in Product
			sodium		
			oxygen		

Na (*s*) and O_2 (*g*) are pure elements. Because the ***total charge of any pure element or compound is always ZERO***, and these two substances are each composed of *only one type of element*, the charge of each of the atoms in a piece of pure sodium metal or a sample of pure oxygen gas ***is equal to ZERO***.

Now let's consider the charges on sodium and oxygen in the **Na_2O** (*s*) ***product***.

Since sodium oxide (Na_2O) is a ***compound***, it has a total charge = ZERO. We ***also*** must recognize that sodium oxide is an ***ionic compound*** (we know this because there is a metal in the compound).

Although the *total charge* of Na_2O is ZERO, the sodium cations and oxygen anions are charged particles; the charge of a sodium ***ion*** is always 1+, the charge of an oxide ion is always 2-.

- Note that ions combine in a ratio such that the total charge of a compound = ZERO; that is why sodium oxide has the formula Na_2O.

Now we will examine how electrons are transferred from one reactant to the other in the reaction.

We will do so by considering what happens to the charge of sodium and oxygen as they are converted from reactants to products.

- As a reactant, sodium exists as Na (*s*) and has a charge of ZERO.
- In the product (Na_2O), ***each*** sodium has a charge of **1+**.
- Since the charge of sodium _____________ in the reaction, we can conclude that *each* sodium must have __________ ***one electron***:

$$Na^o \rightarrow Na^+ + e^-$$

When a species loses electron(s), we call that ***oxidation***. ***Sodium was oxidized in this reaction***.

An oxidation cannot occur without a reduction; the electron(s) will be transferred to another atom or ion.

Where did the electron that sodium lost go?

Since oxygen is the only other reactant, it must have been transferred to the oxygen! Let's verify this by looking for a change in the charge of oxygen in the reaction process.

- In the reactant, oxygen exists as O_2 and has a charge of ZERO.
- In the product (Na_2O), oxygen has a charge of 2-.

Because the charge of oxygen ______________ *(by two)* in the reaction, we can conclude that ***each*** oxygen (in O_2) must have ___________ ***(two) electrons***:

$$O^o + 2\,e^- \rightarrow O^{2-}$$

When a species ***gains*** electron(s), we call that *reduction*. ***Oxygen was reduced in this reaction***.

The fact that ***each*** oxygen gains ***two*** electrons and ***each*** sodium only loses ***one*** electron is accounted for in the balanced chemical equation.

- ***Four sodium atoms*** react with ***two oxygen atoms*** (in O_2). The ratio of sodium atoms to oxygen atoms is 2:1. Two sodium atoms are required to reduce ***each*** oxygen atom.

Example Problem

Answer the questions for the reaction of magnesium metal and chlorine gas.

$$Mg\,(s) + Cl_2\,(g) \rightarrow MgCl_2\,(s)$$

a) What is the charge of each of the magnesium atoms in the reactant [Mg (*s*)]? **0 (ZERO)**

b) What is the charge of each of the magnesium ions in the product? **2+**

c) Did each magnesium atom ***gain*** or ***lose*** electron(s) in this reaction? If so, how many? **lost 2*e*⁻**

$$Mg^o \rightarrow Mg^{2+}$$

d) Was magnesium oxidized or reduced? **oxidized**

e) What is the charge of each of the chlorine atoms in the ***reactant*** [Cl_2 (*g*)]? **0 (ZERO)**

f) What is the charge of the ***each*** of the chloride ions in the product? **1-**

g) Did each chlorine atom gain or lose electron(s) in this reaction? If so, how many? **gained 1e^-**

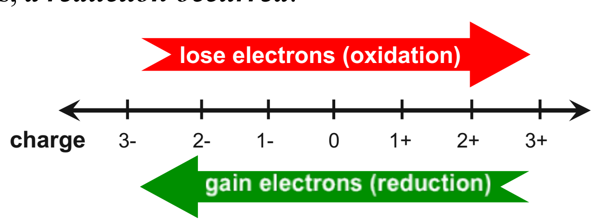
$$Cl^o + e- \rightarrow Cl^-$$

h) Was chlorine oxidized or reduced? **reduced**

Summary of Redox Reactions of *Inorganic* Compounds

It is possible to identify redox reactions for inorganic compounds by inspecting the chemical equation and determining if electrons are *transferred from one species to another*.

- If the **charge** of an atom or ion in a reactant was ***increased*** (toward positive) in the conversion of reactants to products, ***an oxidation occurred***.
- If the **charge** of an atom or ion in a reactant was ***decreased*** (toward negative) in the conversion of reactants to products, ***a reduction occurred***.



Understanding Check

Answer the questions that follow for the reaction of lithium metal and bromine gas:

$$2\ Li\ (s)\ +\ Br_2\ (g)\ \rightarrow\ 2\ LiBr\ (s)$$

a. What is the charge of each of the lithium atoms in the reactant [Li (*s*)]?

b. What is the charge of each of the lithium ions in the product?

c. Did each lithium ***gain*** or ***lose*** electron(s) in this reaction? If so, how many?

d. Was lithium oxidized or reduced?

e. What is the charge of each of the bromine atoms in the ***reactant*** [Br_2 (*g*)]?

f. What is the charge of each of the bromide ions in the product?

g. Did each bromine gain or lose electron(s) in this reaction? If so, how many?

h. Was bromine oxidized or reduced?

Redox Reactions of *Covalent* Compounds

We identified redox reactions for *inorganic compounds* by inspecting the chemical equation and determining if there was a change in the charge of atoms or ions. The transfer of electrons to or from covalent compounds is *not as easily recognized* as was the case for elements and ionic compounds in inorganic redox reactions.

For ***covalent compounds***, such as organic and biological compounds, the gaining and losing of electrons is the result of a gain or loss of bond(s) to ____________ ***atoms*** or ___________ ***atoms***.

For our purposes, oxidation and reduction for covalent compounds can be identified as follows:

- An atom in a covalent compound is **oxidized** if it ***gains*** *bond(s)* to ***oxygen*** and/or ***loses*** *bond(s)* to ***hydrogen***.
- An atom in a covalent compound is **reduced** if it ***loses*** *bond(s)* to ***oxygen*** and/or ***gains*** *bond(s)* to ***hydrogen.***

An example of a redox reaction for *covalent compounds* is the **combustion of hydrocarbons**.

- In the "complete combustion" of hydrocarbons, the hydrocarbon molecules react with oxygen gas (O_2) to form carbon dioxide and H_2O vapor. When there is an insufficient supply of O_2, "incomplete combustion" occurs and other products, such as carbon monoxide, are formed. In this course, when the term "combustion" is used, we will consider that to mean "complete combustion."

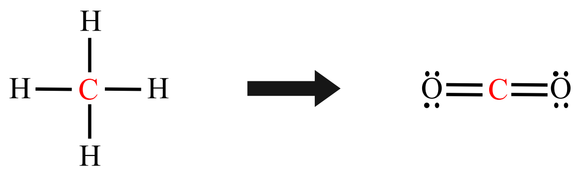
A specific example of the combustion of a hydrocarbon is the reaction of methane and oxygen gas:

$$CH_4\ (g)\ +\ \mathbf{2}\ O_2\ (g)\ \rightarrow\ CO_2\ (g)\ +\ \mathbf{2}\ H_2O\ (g)$$

$$CH_4\,(g) \;+\; \mathbf{2}\;O_2\,(g) \;\rightarrow\; CO_2\,(g) \;+\; \mathbf{2}\;H_2O\,(g)$$

Let's use our criteria for oxidation and reduction of covalent compounds to determine which species was oxidized and which was reduced in this reaction.

- Carbon is bonded to four hydrogen atoms and to zero oxygen atoms as a reactant (in CH_4).
- After reacting, carbon is bonded to zero hydrogens and two oxygens.



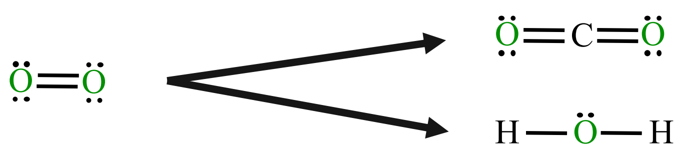
This change in carbon's bonding matches our criteria for the ***oxidation*** of a covalent compound.

An atom in an covalent compound is **oxidized** if it ***gains*** *bond(s)* to ***oxygen*** and/or ***loses*** *bond(s)* to ***hydrogen***.

The carbon in methane was **oxidized** in this reaction.

Let's consider *oxygen* next. *Each* oxygen atom in O_2 is bonded to ***one*** **other** *oxygen atom* and **zero** *hydrogen atoms* before reacting.

After the reaction, oxygen appears in *both* products.



- The oxygen in CO_2 is bonded to zero *other* oxygen atoms (it ***lost*** a bond to oxygen in the reaction).
- The oxygen in H_2O is bonded to zero *other* oxygen atoms and to two hydrogen atoms (it ***lost*** a bond to oxygen **and *gained*** two bonds to hydrogen).

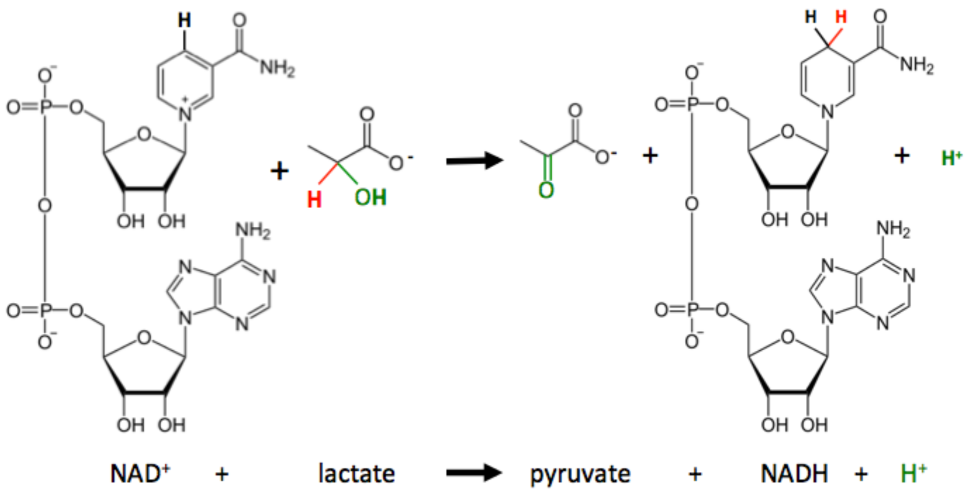
All of these changes in oxygen's bonding match our criteria for the ***reduction*** of a covalent compound.

An atom in a covalent compound is **reduced** if it ***loses*** *bond(s)* to ***oxygen*** and/or ***gains*** *bond(s)* to ***hydrogen.***

The oxygens in $\mathbf{O_2}$ **(*g*)** were **reduced** in this reaction.

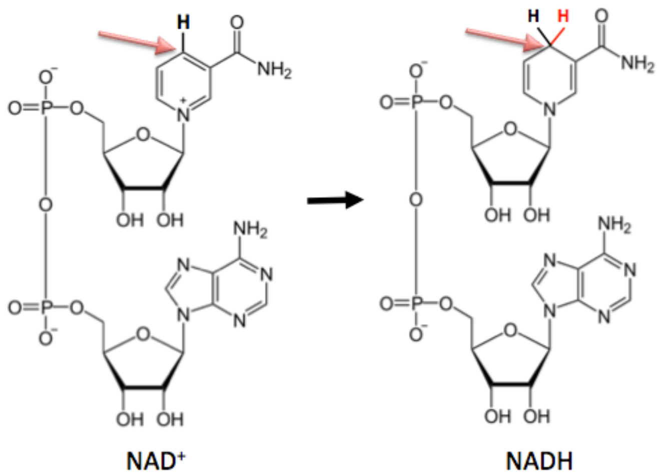
An example of a redox reaction that is involved in biological systems is the reduction of NAD^+/oxidation of lactate.

- This is an important reaction that's involved in the metabolism of food.



We can use our criteria for oxidation and reduction of organic compounds to determine which species was oxidized and which was reduced in this reaction.

Let's consider the change that occurred for **NAD^+**.



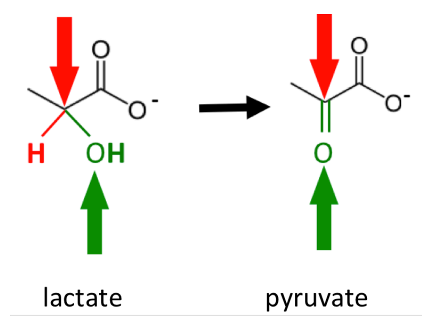
Look at the carbon *indicated by the arrow* at the very top of the six membered ring of the NAD^+. This carbon is bonded to ***one*** hydrogen atom as a reactant (in NAD^+). After reacting, this same carbon is bonded to ***two*** hydrogens.

This change in that carbon's bonding matches our criteria for the ***reduction*** of an organic compound.

An atom in a covalent compound is **reduced** if it ***loses*** *bond(s)* to ***oxygen*** and/or ***gains*** *bond(s)* to ***hydrogen.***

This carbon in NAD^+ was *reduced* in this reaction. We often alternatively state this as, "NAD^+ was *reduced*."

Now, let's consider the change that occurred for *lactate*.



Both a carbon *<u>and</u>* an oxygen in lactate ***lost bonds*** to hydrogen.

Both of these changes match our criteria for the ***oxidation*** of an organic compound.

An atom in a covalent compound is **oxidized** if it ***gains*** *bond(s)* to ***oxygen*** and/or ***loses*** *bond(s)* to ***hydrogen***.

A carbon and an oxygen in lactate were *oxidized* in this reaction. We often alternatively state this as, "lactate was *oxidized*."

Understanding Check

Indicate which atoms in the reactants in each of the following reactions were ***oxidized*** **and** which were ***reduced***.

a. $2\ H_2\,(g) + O_2\,(g) \rightarrow 2\ H_2O\,(g)$

b. ethene + $H_2\,(g)$ $\rightarrow$ ethane

Some Industrial Uses of Redox Reactions

It is estimated that 50% of the world's population eat food that is grown using ***nitrogen fertilizers*** [1].

- The fertilizer industry employs the *reduction* of N_2 (from air) and the *oxidation* of H_2 in order to make ammonia (NH_3) in a process called **nitrogen fixation**.

$$N_2 + 3\ H_2 \rightarrow 2\ NH_3$$

- Ammonia is either used directly or as a starting material for the synthesis of other nitrogenous fertilizers.

Many metals are mined as cations in ionic compounds (metal oxides, metal sulfides, metal carbonates, or metals combined with other anions) and then the metal cations are ***reduced*** to pure metals in an industrial process called **smelting**.

- For example, iron(II) oxide, copper nitrate, and lead sulfide, are converted to pure iron, copper, and lead, respectively. The reducing agent is usually carbon monoxide.

Electroplating processes use redox reactions to apply a thin coat of metal on another metal or conductive material.

- You may have heard the term "gold plated."
- In *gold plating*, silver or copper is often used as the *base metal*. The base metal is submerged in water containing dissolved gold cyanide (AuCN). As electrons are supplied to the base metal from the power source, they are transferred to the gold ions (reduction) and pure gold metal is deposited/plated on the base metal's surface.
- This process is used in the jewelry business and is also an important step in the manufacture of electronic devices such the electrical connectors on microchips.

[1]*Erisman, Jan Willem; MA Sutton, J Galloway, Z Klimont, W Winiwarter (October 2008). "How a century of ammonia synthesis changed the world", Nature Geoscience* ***1*** *(10): 636.*

Cathodic corrosion protection is a process that combines a small quantity of a more easily oxidized metal (called the sacrificial metal) with the metal that is to be *protected* from oxidation.

- Metal oxidation is also called **corrosion**.
- In cathodic protection, the sacrificial metal corrodes before the protected metal.
- An example of cathodic protection is ***galvanized*** steel.
- In the galvanization process, a thin layer of zinc is applied to steel (iron) as a sacrificial metal.

The process called "**bleaching**" uses fairly strong oxidizing agents.

- Oxidizing agents accept electrons from the species that is being oxidized.
- The substance that is commonly referred to as liquid bleach uses the hypochlorite ion (ClO^-) as an oxidizing agent.
- Bleach removes electrons from the molecules that make up "stains" or infectious agents.
  - In the case of stains, when oxidized by bleach, they no longer absorb visible light and therefore can no longer be seen.
- Other bleaching methods use oxidizing agents such as hydrogen peroxide (H_2O_2) instead of hypochlorite.

Biochemical Redox Reactions

You will learn about the chemistry of biological systems (biochemistry) in the later chapters. *You will see many redox reactions occurring in biology.*

Biological systems employ a series of chemical reactions in order to get energy from food molecules in the ***cellular respiration*** process.

- Many of these reactions are ***redox reactions***.

The energy in food molecules originates from the sun. Sunlight is converted to chemical energy (in the form of carbohydrate molecules) in a series of chemical reactions called ***photosynthesis***.

- Many of the reactions of ***photosynthesis*** are ***redox reactions***.

Summary of Redox Reactions

Oxidation is the ***loss of electron(s)*** by an atom, ion, or molecule in a chemical reaction.

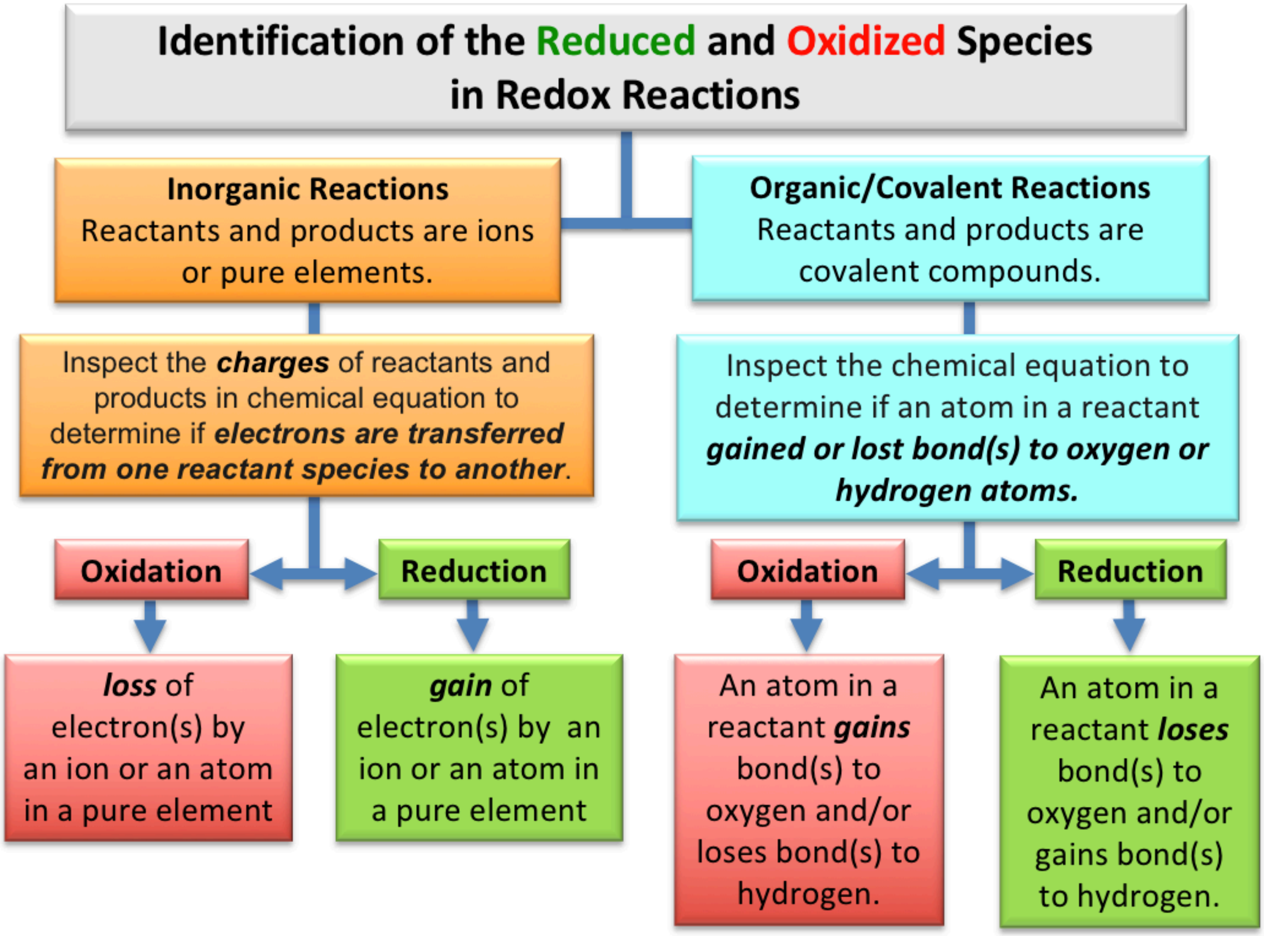
Reduction is the ***gain of electrons*** by an atom, ion, or molecule in a chemical reaction.

It is possible to identify ***redox reactions*** for **inorganic compounds** by inspecting the chemical equation and determining if electrons are transferred from one species to another.

- One atom or ion in a **reactant** will **lose electrons** and therefore its ***charge*** will **increase** (in the positive direction) by one charge unit for every electron that is lost.
- The ***charge*** of the atom or ion in the **reactant** that **gains** the electron(s) will **decrease** (toward negative values) by one charge unit for every electron that is gained.

For **covalent compounds** such as organic molecules:

- An atom in a covalent compound is **oxidized** if it ***gains*** *bond(s)* to ***oxygen*** and/or ***loses*** *bond(s)* to ***hydrogen***.
- An atom in a covalent compound is **reduced** if it ***loses*** *bond(s)* to ***oxygen*** and/or ***gains*** *bond(s)* to ***hydrogen.***



Redox Reaction Terminology

The ***reactant that was oxidized*** is sometimes referred to as the "**reducing agent**" because it *transferred its electron(s)* to the reactant that was reduced.

Conversely, the ***reactant that was reduced*** is sometimes referred to as the "**oxidizing agent**" because it *accepted electron(s)*from the reactant that was ***oxidized***.

Reactions of Organic Molecules

In this remainder of this chapter, I will discuss ***four classes of organic reactions*** that involve the ***families*** of organic molecules that you were introduced to in chapter 4: hydrocarbons, alcohols, carboxylic acids, and esters.

Before we begin, let's review the structure of alcohols, carboxylic acids, and esters.

Review of *families* of Organic Molecules

The Alcohol Family

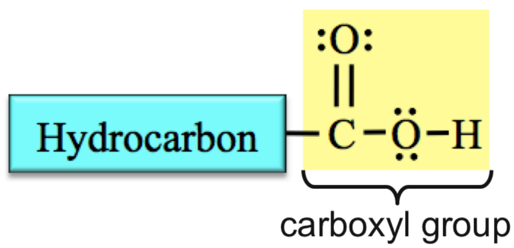
Alcohols contains one or more _______________ **(-OH)** functional groups attached to a ***hydrocarbon***.

The general form of an alcohol is shown below with the hydroxyl group highlighted in yellow.

The Carboxylic Acid Family

Carboxylic acids contain a _______________ ***functional*** group attached to a hydrocarbon.

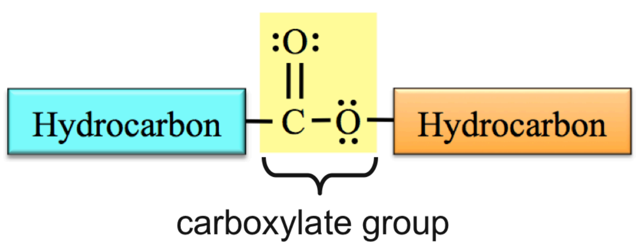
Carboxyl groups contain both a ***carbonyl group***, which is a carbon double bonded to an oxygen (**C=O**), ***and*** a ***hydroxyl group*** **(-OH)** that are connected to each other and the hydrocarbon (alkyl group) part.



The Ester Family

Esters contain a ___________________ ***functional*** group that is bonded *between* ***two*** hydrocarbon parts.

Carboxylate groups contain both a ***carbonyl group*** (**C=O**), ***and*** an ***oxygen atom.***



The boxes that represent hydrocarbons (alkyl groups) are shaded with different colors because the two hydrocarbon parts *are not always identical*.

Reactions of Organic Molecules

The ***four classes of organic reactions*** to be studied in this section are:

1. **Hydrogenation: Reduction of Alkenes**
2. **Hydrolysis of Esters**
3. **Hydration of Alkenes**
4. **Dehydration of Alcohols**

The **educational goal** for this section is, ***if you are given the specific reactant(s) for any of these four classes of reactions, you should be able to predict (draw) the product(s).***

A good way to do this is to know the "**general form**" of the organic reaction.

I will elaborate on what is meant by the "***general form***" of an organic reaction when I discuss the hydrogenation (the reduction of alkenes).

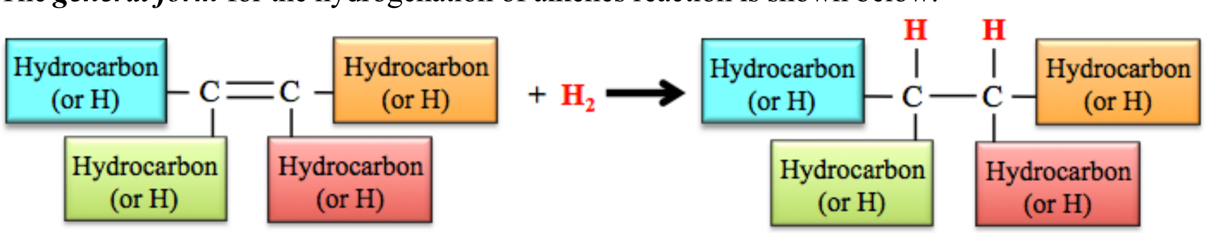
1) Hydrogenation: Reduction of Alkenes

Alkenes and other unsaturated hydrocarbons react with hydrogen gas (**H_2**) in a reaction called ________________.

In the presence of a *catalyst*, such as platinum, a hydrogen atom from **H_2** is added to each of the double bonded carbon atoms in the ***alkene*** to produce the corresponding ______________.

- Hydrogen gas acts as a *reducing agent*; the carbon atoms in an alkene are ***reduced*** (*they gain hydrogen atoms)*.

The ***general form*** for the hydrogenation of alkenes reaction is shown below.

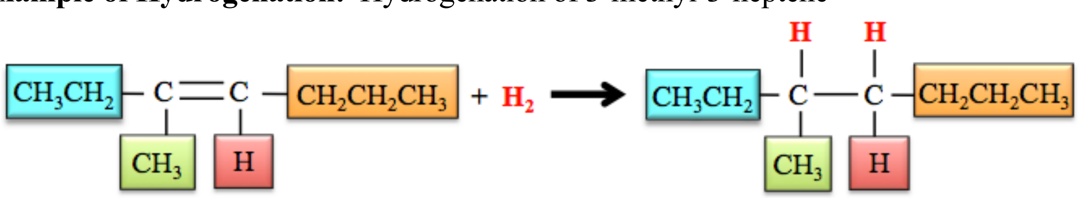


Alkene **Alkane**

The ***general form*** of an equation contains the generic structures of products and reactants.

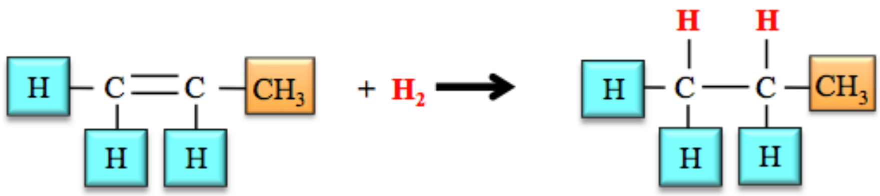
- For example, in the general form for the hydrogenation equation above, a *generic structure* representing ***any alkene*** is drawn as the *reactant* and a *generic structure* representing ***any*** alkane is drawn for the *product*.
- The boxes represent any hydrocarbon (alkyl group) or a hydrogen atom, and are shaded with different colors (in the video) to indicate that ***the hydrocarbon parts may or may not be identical***.

Specific Example of Hydrogenation: Hydrogenation of 3-methyl-3-heptene



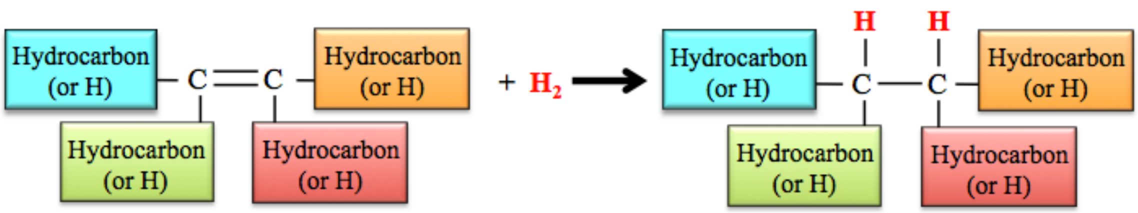
3-methyl-3-heptene **3-methylheptane**

Specific Example of Hydrogenation: Hydrogenation of propene



propene **propane**

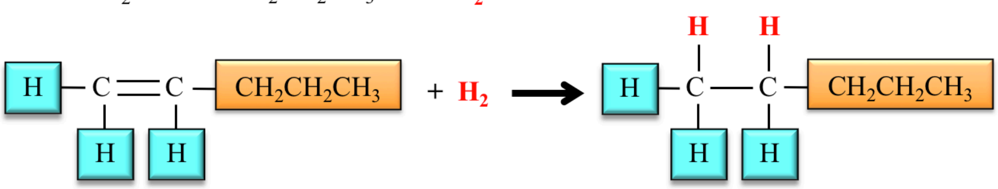
Knowing the "***general form***" of an organic reaction allows you to predict and draw the product(s) when given specific reactant(s).



Alkene **Alkane**

Example: Draw and name the ***product*** of the following reaction:

$CH_2{=}CHCH_2CH_2CH_3 + H_2 \longrightarrow$



$CH_2{=}CH\,CH_2CH_2CH_3 + H_2 \longrightarrow CH_3CH_2CH_2CH_2CH_3$

1-pentene **pentane**

Before I ask you to try a couple of problems on your own, I want to show you another way to predict the product for hydrogenation of alkene reactions. This method will be useful in several other types of reactions that you will learn.

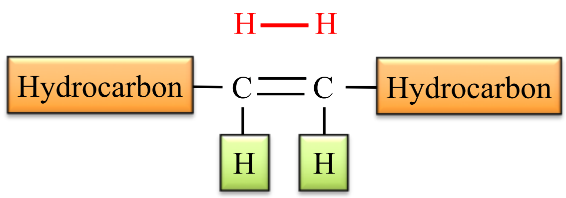
In the hydrogenation reaction, we made *new bonds* to ***both of the atoms (carbons)*** that were connected to each other with a double bond.

Chemical reactions where new bonds are formed to atoms at each end of a double bond occur so frequently that organic chemist have a special name for it: "***addition across a double bond***."

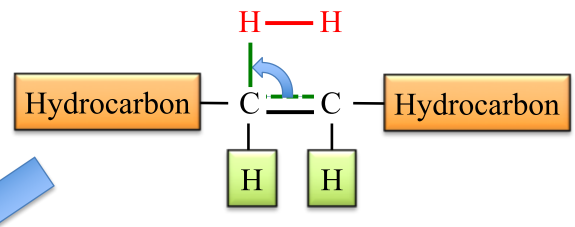
Products for reactions where *addition across a double bond* occurs can be easily predicted by learning the following method of "***flipping***" bonds.

Addition Across a Double Bond

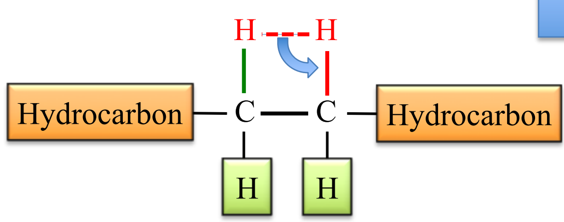
Step #1: Draw the **H-H** directly above the double bond.



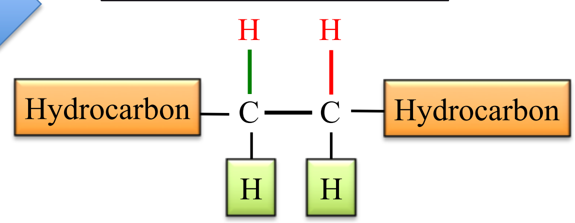
Step #2: Flip *one* of the bonds from the double bond **up** to the **H**



Step #3: Flip the **H-H** single bond **down** to the **C**



The product is an alkane



Understanding Check

Draw and name the product formed when each of the alkenes listed below react with H_2.

a. 1-butene

b. *cis*-2-butene

c. *trans*-2-pentene

Water is a reactant or product in a number of reactions important to organic and biochemistry.

The next three classes of organic reactions involve a water molecule as either a reactant or product.

2) Hydrolysis of Esters

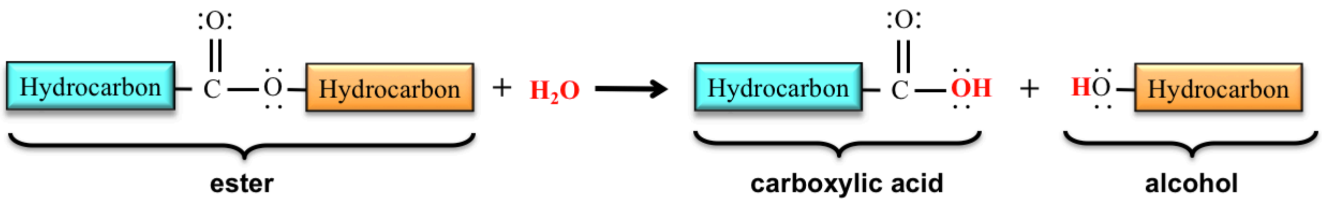
In a _____________ reaction, water (hydro) is used to break (lyse) a bond in a molecule.

The **hydrolysis of esters** reaction occurs in nature and has industrial uses.

- Example: Triglycerides (fats) are esters and undergo hydrolysis reactions in the digestion process.

In the hydrolysis of an ester, a water molecule breaks a bond in the ester to form a _______________ _____ and an _________________.

- The general form of the hydrolysis of esters reaction is shown below:



Knowing the "*general form*" of the hydrolysis of esters, and then identifying the hydrocarbon parts will allow you to predict and draw the carboxylic acid and alcohol products when given a specific *ester* reactant.

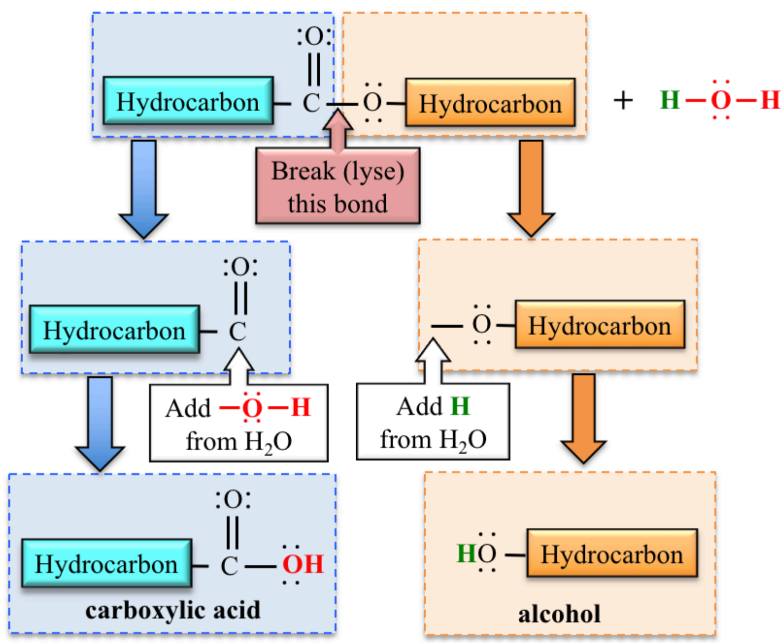
I want to show you *another way* to predict the products for hydrolysis of ester reactions.

Step 1: Draw the structural formula of the ester and identify the hydrocarbon parts.

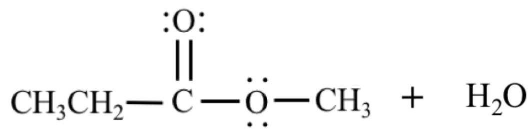
Step 2: Break (lyse) the carbon-oxygen ***single bond*** between the *carbonyl* carbon and the oxygen. The *carbonyl* carbon is the carbon that is double bonded to an oxygen.

Step 3: Add the **–OH** from the water to the *carbonyl* carbon and then add the **H** from the water to the oxygen on the ***other fragment***.

When you use these three steps, it will result in the correct ***carboxylic acid*** and ***alcohol.***



Example: Predict the products of the hydrolysis of the ester shown here.

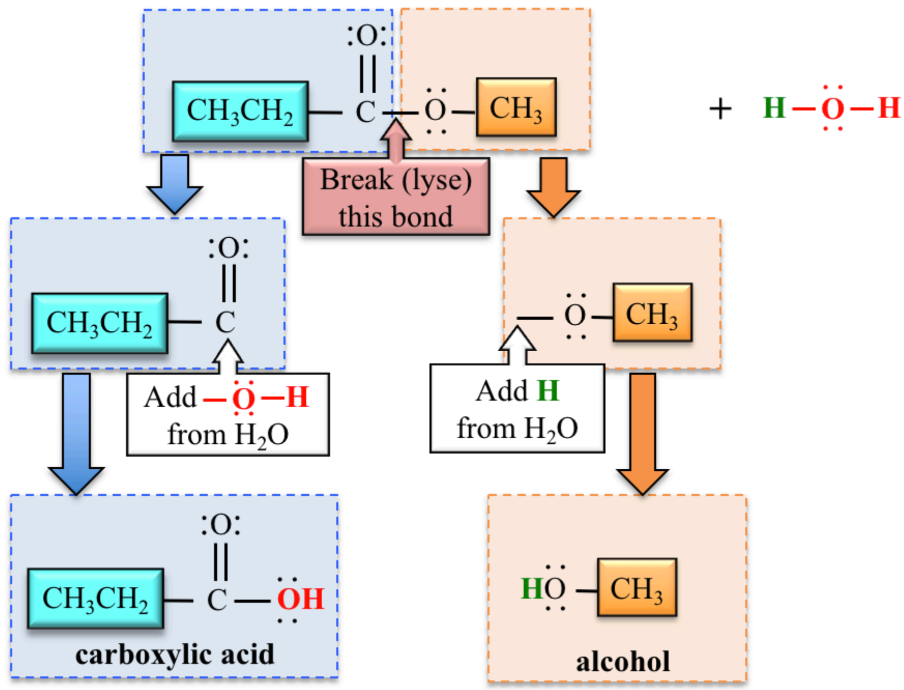


Solution:

Step 1: Draw the structural formula of the ester and identify the hydrocarbon parts.

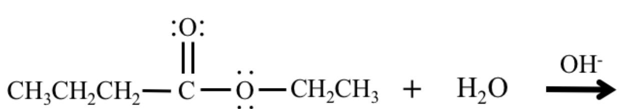
Step 2: Break (lyse) the carbon-oxygen ***single bond*** between the *carbonyl* carbon and the oxygen. The *carbonyl* carbon is the carbon that is double bonded to an oxygen.

Step 3: Add the **–OH** from the water to the *carbonyl* carbon and then add the **H** from the water to the oxygen on the ***other fragment***.



Understanding Check

Draw the structural formulas of **both products** in the hydrolysis of ethyl butanoate:

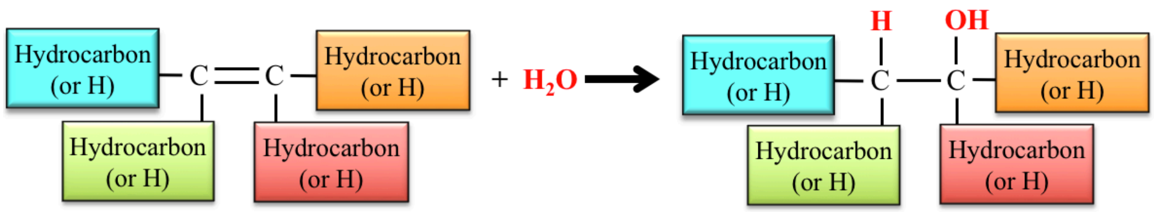


ethyl butanoate

3) Hydration of Alkenes

In _______________ of alkene reactions, alkenes react with water molecules to produce alcohols.

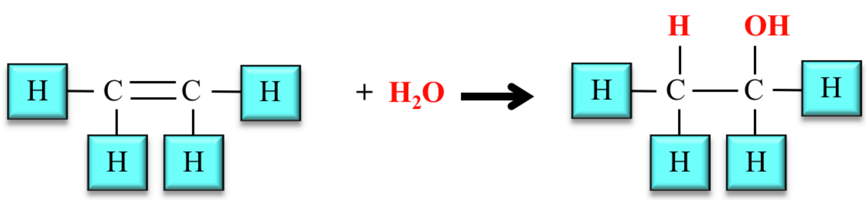
The ***general form*** for the hydration of an alkene reaction is shown below:



Alkene **Alcohol**

- A hydrogen atom from **H_2O** is added to one of the double bonded carbon atoms and the **–OH** from the **H_2O** is added to the ***other*** double bonded carbon atom in the ***alkene*** to produce the corresponding ***alcohol***.
- The **double bond** in the *alkene* is converted to a **single bond** in the *alcohol*.

A specific example of the hydration of alkenes reaction is the reaction of ***ethene*** and **H_2O**:



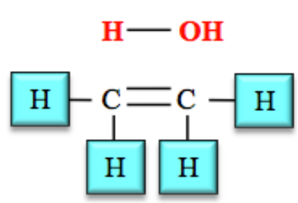
You may have noticed the similarity of this *hydration reaction* to the *hydrogenation of alkenes*.

Hydration reactions are another example of reactions in which a molecule is ***added across a double bond***.

In the *hydration of alkenes reaction*, water is added ***across the double bond of an alkene,*** therefore we can use the *bond flipping method* to predict the structure of the alcohol that is produced.

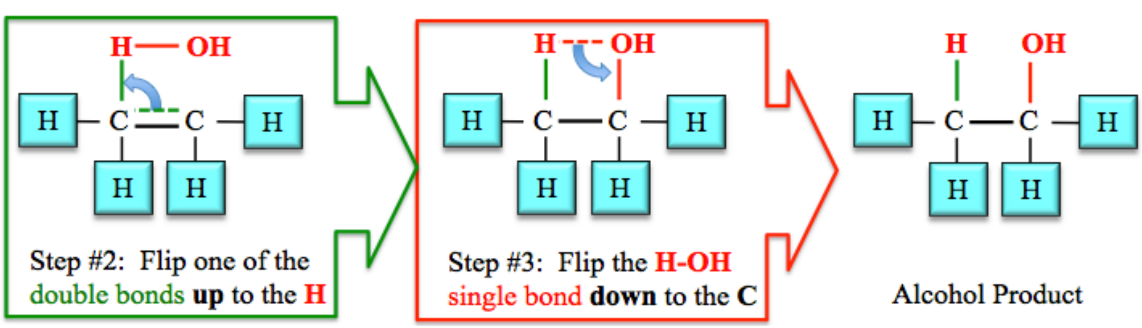
EXAMPLE: ***Addition of H_2O across a double bond***. Add **H_2O** across the double bond of **ethene** (hydration of ethene).

Step 1: Draw the molecule to be added across the double bond and the molecule with the double bond as shown here:



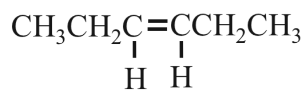
Step #1:
Draw the **H - OH** directly above the double bond.

Steps 2 and 3: ***Flip*** the bonds as shown below to get the product of the reaction.



Understanding Check

Predict the product formed by the hydration of this alkene:



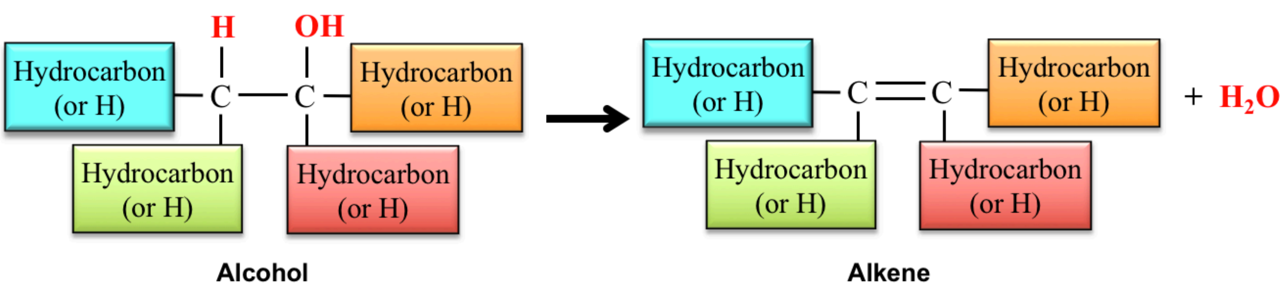
4) Dehydration of Alcohols

Dehydration of alcohols is the ***reverse*** of hydration of alkenes.

H_2O is ______________ from an alcohol to form an alkene.

- A hydroxyl group (**-OH**) is removed from a carbon atom and an **H** is removed from a carbon that is __________________ to the carbon that was bonded to the hydroxyl group. A double bond forms between these two carbons.

The ***general form*** for the *dehydration of an alcohol reaction* is shown below:



Since this reaction is the ***reverse** of the hydration of alkenes* reaction, we can ***flip*** the bonds in the ***opposite order*** to that which we used when we added water *across the alkene double bond*.

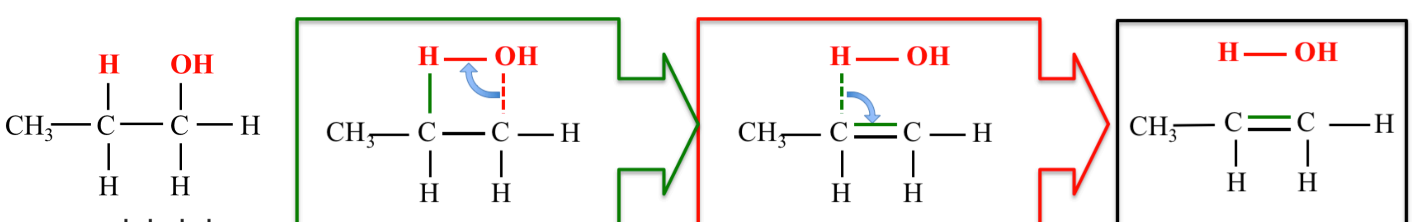
Doing so can be very helpful in determining the alkene product in a dehydration of alcohol reaction.

I will use the dehydration of *propyl alcohol* as an example. Beginning with the structure of the alcohol, perform the following two steps:

Step #1: Flip the bond between the C and **OH** **up** to an **H** on *an adjacent* carbon.

Step #2: Flip the **C--H** bond from the *adjacent* carbon **down** to make a double bond to the C that originally carried the OH

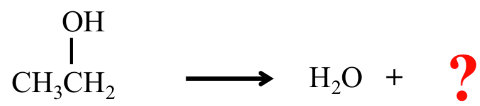
Alkene and **H_2O** Products



propyl alcohol

Understanding Check

Name and draw the *condensed structural formula* for the alkene that is produced when ethyl alcohol undergoes a dehydration reaction:



Summary of Classes of Organic Reactions

Organic Reaction Class	Reactant(s)	Product(s)
Hydrogenation of Alkenes	Alkene + H_2	Alkane
Hydrolysis of Esters	Ester + H_2O	Carboxylate ion + Alcohol
Dehydration of Alcohols	Alcohol	Alkene + H_2O
Hydration of Alkenes	Alkene + H_2O	Alcohol

Recommendation: Work on the "Chapter 6 Reactions Worksheet" as soon as you can.

- You can find this worksheet in the lecture notes package or on the website.