




You Be The
Chemist
CHALLENGE®

PASSPORT TO SCIENCE EXPLORATION
CHEMISTRY CONNECTIONS



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Welcome to the *You Be The Chemist Challenge*®!

Welcome to the *You Be The Chemist Challenge*® (Challenge)! The Challenge is an exciting academic competition that will expand your knowledge of the science of chemistry and beyond.

The information in this *Passport to Science Exploration* (Passport) exposes you to the fascinating world of chemistry as it applies to your everyday life. It shows you how chemicals shape your world. Read and review the information in the Passport, including all relevant examples, to further your understanding of chemistry. Challenge competition questions are based on an overall understanding of chemistry concepts and the relationships among chemistry concepts. Challenge questions are largely derived from information contained in CEF's study materials but are not limited solely to this information. We encourage you to seek additional examples and explanations of chemistry concepts. Doing so will help you answer questions that require you to connect concepts and apply your knowledge of chemistry to both familiar and unfamiliar situations.

How do I use the Passport?

The information in the Passport is supplied to help you succeed at every level of the Challenge. The Passport is divided into three sections that correspond to the different levels of the Challenge – local, state, and national.


This portion, *Chemistry Connections*, takes you a step further as you explore the world of chemistry. You need to be familiar with this information and the *The Core of Chemistry* portion for State Challenge competitions¹, as well as for the National Challenge. If you are not sure about how to prepare for a particular level of competition, ask your teacher or Local Challenge Organizer. You may also contact the Chemical Educational Foundation® (CEF) at challenge@chemed.org.

Again, you should begin your studies with *The Core of Chemistry* and *Chemistry Connections* for your State Challenge competition. However, we also encourage you to explore *all* the study materials provided on CEF's website, www.chemed.org, as well as outside resources. The more you explore, the more likely you are to find answers to the questions you have about the world around you!

The tips below will help you get the most out of the information provided.

1. Read over the Table of Contents first. This introduces you to the concepts covered.
2. Review the Objectives listed at the beginning of each section to become familiar with the topics you will learn. Once you have completed your study of the entire section, go back and review the Objectives again. Can you complete the tasks listed in the Objectives? If not, review the section again.
3. Focus on the definitions of the **bolded terms**. Then, follow up by reading the bulleted information.
4. Use the diagrams, pictures, and illustrations to gain a better understanding of the concepts.

¹The categories of CEF study materials and/or specific concepts covered in Challenge competitions may be adjusted at the discretion of CEF or of Challenge Organizers (with the approval of CEF). Students will be notified by their Local/State Challenge Organizer in the event of such changes.

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5. Read through material labeled as a “Quick Fact” (in circles). Quick Facts are not necessary to understanding the major concepts. However, they give further details and provide applications to help you understand the material even better. Although the Quick Facts are not necessary for understanding the basic concepts, they are still valid competition material. Be sure to review them, too!
 6. Read the information in the History boxes. These boxes provide a variety of background knowledge about chemistry concepts. These boxes may highlight a famous scientist who made an important discovery related to a concept. They may also explain previous beliefs and how those beliefs have changed as we have learned more about our world. Either way, the History boxes help to connect science of the past to science today.
 7. Use the information in the Element boxes to learn more about specific elements within a section.
 8. Read through material in the circles labeled “Think About It.” These present questions related to the material in a particular section. The answers to some Think About It questions may be obvious after reading the material. However, some answers may not even be known to scientists. These questions are placed in the Passport to make you think! Don’t worry if the answer to a Think About It question is not obvious to you. Use these questions to explore more about chemistry and find out what questions scientists have or have not been able to answer. (Search the Internet, check chemistry books, or ask a scientist or teacher to find an answer.)

Once you are finished with a section, do a quick review to make sure you learned all the concepts introduced in that section. It might help to make flashcards, too. If you find that you still do not understand something, pull out a science textbook and look in the index for the information or conduct a search on the Internet (be sure to find a reliable source). If the explanation you find is unclear, ask your science teacher for help.

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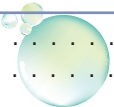
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SECTION I: CHEMICAL FORMULAS

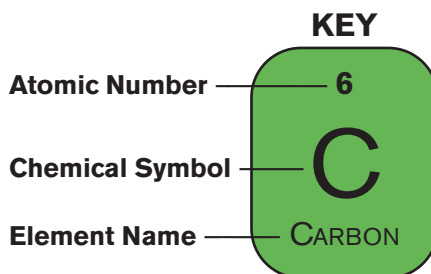
OBJECTIVES

- Identify the chemical symbols for different elements.
- Identify and write chemical formulas using chemical symbols.
- Recognize common chemical compounds and their formulas.

CHEMICAL SYMBOLS

As you already know, **chemical symbols** are used to represent the elements. Each element has its own symbol that is different from all other chemical symbols. These symbols are made up of one or two letters (except for some new, unconfirmed elements). The first letter of a chemical symbol is always capitalized. If a chemical symbol has a second letter, it is written in lowercase. For example, the chemical symbol for oxygen is O. The chemical symbol for calcium is Ca.

Chemical symbols are used on the periodic table (as shown below).



WRITING CHEMICAL FORMULAS

Chemical symbols represent elements. When elements combine to form compounds, scientists use chemical symbols to write these combinations. A **chemical formula** is the way scientists write a compound using chemical symbols.

EXAMPLE:

Water is a compound that contains the elements hydrogen and oxygen. Two hydrogen atoms and one oxygen atom combine to form water.

The chemical symbol for hydrogen is H. The chemical symbol for oxygen is O. Therefore, the chemical formula for water is H₂O.

Chemical formulas use **subscripts** to indicate how many atoms of each element there are. Subscripts are the numbers written to the lower right of a chemical symbol. Notice that the chemical formula for water has the subscript 2 after the H. Therefore, we know that there are two hydrogen atoms in the compound. There is no subscript after the O, so we know that there is only one oxygen atom in the compound.

When a metal chemically combines with a nonmetal, the chemical symbol for the metallic element is usually written first. Sodium (a metal) and chlorine (a nonmetal) combine to form the compound sodium chloride. Notice that the metal is written first.

Oxidation numbers are also important when writing chemical formulas. An oxidation number can be used to show how many electrons an atom gains, loses, or shares when it chemically combines with another element.

- In general, when the outermost energy level of an atom contains eight valence electrons (like the noble gases), the atom is very stable (see section on **Chemical Bonds**). Therefore, it tends not to gain or lose electrons.
- To reach a more stable state, atoms can gain, lose, or share electrons by forming chemical bonds with other atoms.
- Nonmetal atoms tend to gain electrons. Metal atoms tend to lose electrons.
- A neutral atom that gains electrons acquires a negative oxidation number. A neutral atom that loses electrons acquires a positive oxidation number.

When writing chemical formulas, the name of the element with the positive oxidation number is written first. Therefore, since metals tend to give up electrons, they generally have positive oxidation numbers. Nonmetals generally have negative oxidation numbers. Consider writing the formula for sodium chloride (NaCl) again. Sodium has a positive oxidation number (1^+). Chlorine has a negative oxidation number (1^-). These numbers explain why the metal, sodium, is written first in the compound sodium chloride.

Quick Fact

Notice that chlorine changes to chloride in the compound sodium chloride. In many cases, the name of the second element is changed to end in *-ide* in the compound.

This change results from gaining an electron. It reflects the change in oxidation number.

Quick Fact

Many elements have more than one oxidation number. For example, iron may have an oxidation number of 2^+ or 3^+ .

When writing out the name of a compound, scientists use Roman numerals to show the oxidation number of an element in a compound. Iron (II) has oxidation number 2^+ . Iron (III) has oxidation number 3^+ . The usual oxidation number for chlorine in compounds is simply 1^- .



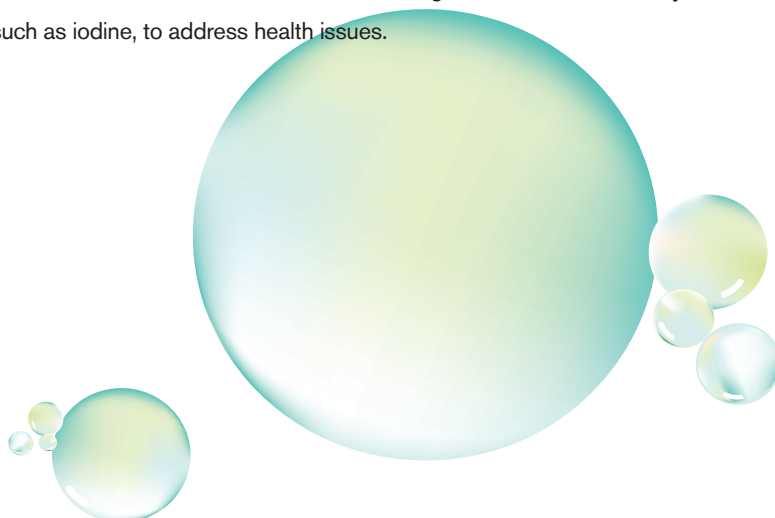
CHEMICAL FORMULAS OF COMMON COMPOUNDS

The following table lists some common compounds and their chemical formulas.

Common Name	Chemical Name	Chemical Formula
Alcohol (grain alcohol)	Ethanol (ethyl alcohol)	C_2H_5OH
Ammonia	Ammonia	NH_3
Bleach (chlorine bleach)	Sodium hypochlorite	$NaOCl$
Chloroform	Trichloromethane	$CHCl_3$
Laughing gas	Nitrous oxide (dinitrogen oxide)	N_2O
Lye	Sodium hydroxide	$NaOH$
Muriatic acid	Hydrochloric acid	$HCl(aq)^*$
Quicklime	Calcium oxide	CaO
Silica (sand)	Silicon dioxide	SiO_2
Rock salt (halite)**	Sodium chloride	$NaCl$
Table sugar (cane sugar)	Sucrose	$C_{12}H_{22}O_{11}$
Vinegar	Acetic acid	$CH_3COOH(aq)^*$
Water	Water	H_2O
Wood alcohol	Methanol (methyl alcohol)	CH_3OH

* An *aqueous solution* is a solution in which the solvent is water. Therefore, vinegar is actually a mixture of acetic acid in water. To indicate an aqueous solution, scientists generally list (aq) after the chemical formula. For example, in the chemical equation $H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$, carbonic acid, H_2CO_3 , is an aqueous solution.

** Halite, commonly known as rock salt, is the mineral form of NaCl. Common table salt is also primarily made of NaCl (generally about 97%-99% NaCl) but may also contain other chemical substances, such as magnesium carbonate. Many brands of table salt also contain other additives, such as iodine, to address health issues.



SECTION II: FORCES OF ATTRACTION

OBJECTIVES

- Explain the forces of gravity and magnetism.
- Describe chemical bonds as a force of attraction.
- Explain the relationship of Coulomb's law and electronegativity to chemical bonding.
- Identify the three primary types of bonds.
- Use Lewis Dot Structures to illustrate bonding.

GRAVITY

When you slide a book over the edge of a table, you expect it to drop to the ground. If you let go of a bag you're holding, you expect the bag to fall to the ground too. Why? Gravity, of course! **Gravity** is the force of attraction between all objects in the universe. It is the force that keeps the planets in our solar system within their orbits around the sun. Otherwise, they would simply float off randomly into space. On the earth, all objects are pulled toward the earth's center. Therefore, objects will fall toward the ground.

Because the force of gravity acts between all objects, any two objects in the universe will be attracted to each other. Gravity pulls you toward the ground, but you are also attracted to all of the other objects around you! So, why does your backpack fall to the ground rather than orbit around your body? The answer has to do with the amount of force that each object exerts.

One factor that affects the gravitational force (gravitational attraction) between objects is mass. Objects with greater mass will have greater gravitational force. Therefore, your backpack will fall toward the earth because the mass of the earth is much, much greater than the mass of your body. Similarly, the mass of the sun is so great that it keeps all the planets in orbit.

Another factor that affects the gravitational attraction between objects is distance. Objects that are farther apart have less gravitational force between them. Therefore, as a spacecraft leaves the earth's atmosphere, its attraction to the earth becomes less and less as it moves farther away.

Quick Fact

Since weight is a measure of the force of gravity, an object's weight varies with the strength of the gravitational force acting on it (see section on **Measurement** from *The Core of Chemistry*.)

HISTORY: ISAAC NEWTON (1643-1727)

Sir Isaac Newton is a famous English scientist known for his work in astronomy, physics, mathematics, and chemistry. As a university student, he developed some important ideas, including his theory of gravity. (He wasn't at school when he developed this idea. The school was closed, but he continued to study on his own!)

When Newton observed a falling apple, he wondered why all objects fall to the ground. He concluded that some force pulls objects toward the center of the earth. This force, which he called gravity, acts throughout the universe and pulls all objects toward each other. He further explained gravity through the following formula for universal gravitation:

$$F_g = G \frac{m_1 m_2}{r^2}$$

where F_g is the gravitational force, m_1 and m_2 are the objects' masses, r is the distance between the two objects, and G is the universal gravitation constant.

Although he formulated the theory of gravity early in his career, Newton did not make his idea public until much later. Likewise, Newton did not publish his famous works, *Philosophiae Naturalis Principia Mathematica* (the *Principia*) or *Opticks*, until years after he had written them. The *Principia* provides explanations on gravity and orbital motion and also presents his theory of fluids. *Opticks* explores the refraction of light by a glass prism. In this work, Newton proposes that white light is made of a mixture of different colored rays.

The first scientific achievement that Newton made public was the invention and construction of the reflecting telescope. The basic design of this telescope is still used to develop large telescopes used today.

Newton is also well known for his three laws of motion that form the basis of the theory of motion.



MAGNETISM

Have you ever noticed that certain objects, like paperclips, nails, or hairpins, will stick to a type of object called a magnet. A magnet may also hold important notes or school papers to your refrigerator at home. A **magnet** is an object that creates a strong magnetic field (an area of magnetic force). Only certain metals produce a magnetic field. These metals include iron, nickel, and cobalt. All magnets are made of these metals, called ferromagnetic metals. Therefore, magnetism is a property of some metals but not all.

The ends of a magnet are called its poles. All magnets have two poles, a north pole (N) and a south pole (S). Two unlike magnetic poles will attract each other. Opposites attract! Two like magnetic poles will repel each other. For example,

Quick Fact

If you hang a magnet by a string, its north pole will turn North. Therefore, the north pole is sometimes called the north-seeking pole. Likewise, the south pole will point South.

the north pole of one magnet will attract the south pole of another magnet. The north poles of two magnets will repel each other (as will the two south poles). Therefore, metals with magnetic properties will be attracted to a magnet. Specifically, the north end of a magnetic metal will be attracted to the south pole of a magnet and vice versa. Metals that do not produce a magnetic field will not be attracted to a magnet.

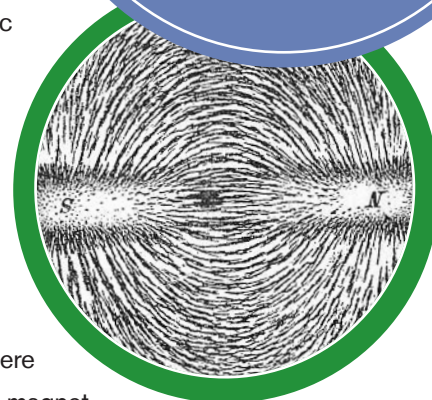
Magnetism is a force of attraction or repulsion between magnetic materials. The magnetic forces are exerted all around the magnet but are the strongest at its poles. The area of magnetic force around a magnet is called a **magnetic field**. This magnetic field causes magnets and magnetic metals to move in certain ways even if they do not touch.

The magnetic field of a magnet is illustrated by *magnetic field lines*. Magnetic field lines spread out from one pole and make a curved path around the magnet to the other pole. These lines do not touch. However, in areas where the lines are closer together, the magnetic force is stronger.

In addition, some magnets have a stronger magnetic force than others. There are two basic kinds of magnets – permanent and temporary. A permanent magnet keeps a certain level of magnetism for a long time. A temporary magnet acts like a permanent magnet when it is within a strong magnetic field. It will lose its magnetism when the magnetic field is removed. For example, any object that is lifted or moved by a magnet acts as a temporary magnet. That object generally loses its magnetism when the permanent magnet is removed. However, in some cases, the magnetized object will still have weak magnetic properties.

Quick Fact

When iron filings are spread around a bar magnet, they will gather around the magnetic field lines. As a result, you are able to “see” the magnetic field surrounding the magnet.



HISTORY: MAGNETIC MATERIALS

Humans have known about magnetic materials for over 2,000 years. The ancient Greeks discovered a unique rock in the city of Magnesia. The rock attracted materials containing iron. The rocks would also attract or repel other similar rocks.

These rocks, known as lodestones, are naturally magnetic rocks that contain the mineral magnetite. Magnetite is an iron oxide with the chemical formula Fe_3O_4 .

Magnetite can be found in many other parts of the world. The Chinese found that if a lodestone was allowed to swing freely from a string, one part of the rock would always point in the same direction. It would point toward the North Star. As a result, the Chinese are credited with inventing the magnetic compass.







Magnetic compasses are still used today. The pointer of a magnetic compass is made with lodestone. Therefore, one end of the pointer will always point toward the North.

CHEMICAL BONDS

Forces of attraction even play an important role inside chemical substances. The hydrogen and oxygen atoms in a molecule of water are held together by certain forces. An intramolecular force or **chemical bond** is the force of attraction that holds together the atoms in a molecule. These bonds are formed as a result of the rearrangement of electrons between the atoms. Therefore, the forces between atoms and molecules are electrical.

The forces of attraction that create chemical bonds are described by **Coulomb's Law**, which tells us: Opposites attract! The charges in an atom or molecule attract if they are different (one positive and one negative). The attraction is greater when the charges are higher.

The charges push each other apart if they are the same (both positive or both negative).

 Correct		The electron (-) in its orbital is attracted to the proton (+) at the center of an atom.*
 Correct		These two electrons repel each other.*
 Incorrect		The two electrons are attracted to a proton, but the electrons are not positioned in the best way.* How should they be positioned?

Atoms chemically combine to form molecules. Therefore, when chemical bonds are formed, broken, or rearranged, a chemical change takes place.

EXAMPLE:

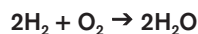
Hydrogen is commonly found on the earth as diatomic hydrogen gas. Two hydrogen atoms combine to make up a molecule of hydrogen gas.

*In the atom images above, the electron cloud has been removed for easier visualization.

When atoms of different elements combine, chemical bonds are formed between the atoms to create a compound.

EXAMPLE:

A molecule of hydrogen gas (H₂) contains two hydrogen atoms chemically bonded together. Likewise, a molecule of oxygen gas (O₂) contains two bonded oxygen atoms. When a molecule of oxygen gas combines with two molecules of hydrogen gas, two molecules of the compound water are formed (see [Chemical Reactions](#) section).



Bonding involves only the electrons in the outermost energy level of an atom (also taking into account sublevels). These electrons are known as the **valence electrons**.

- In general, an atom can hold eight valence electrons. The exceptions are hydrogen and helium.
- This “rule of eight” applies to the electron sublevels. For example, energy level three can hold a maximum of 18 electrons (see section on [Electron Configuration](#) from *The Core of Chemistry*). However, the sublevels determine that there are only a maximum of eight valence electrons that participate in chemical bonding.
- Atoms that have less than eight valence electrons will tend to form bonds with other atoms. They will give, take, or share electrons to reach a stable state with eight valence electrons.

Quick Fact

The first energy level of an atom can hold up to two electrons. Hydrogen and helium are the only elements with electrons that **only** occupy this energy level. The electrons of all other elements occupy additional energy levels. Beyond the first energy level, eight becomes the “magic” number for electrons.

ELECTRONEGATIVITY

Electronegativity is a measure of how strongly the nucleus of an atom attracts electrons in a chemical bond. In other words, electronegativity measures how greedy an atom is for electrons.

- Electronegativity increases from left to right across the periodic table.
 - The number of protons in an atom increases from left to right across the periodic table.
 - Protons have a positive charge, so atoms with many protons have a high nuclear charge.

EXAMPLE:

The electronegativity (from least to greatest) for the second row of the periodic table is: Li, Be, B, C, N, O, F.

Notice that neon (Ne) is not listed. Neon is already full of energy-level 2 (shell 2) valence electrons.

EXAMPLE:

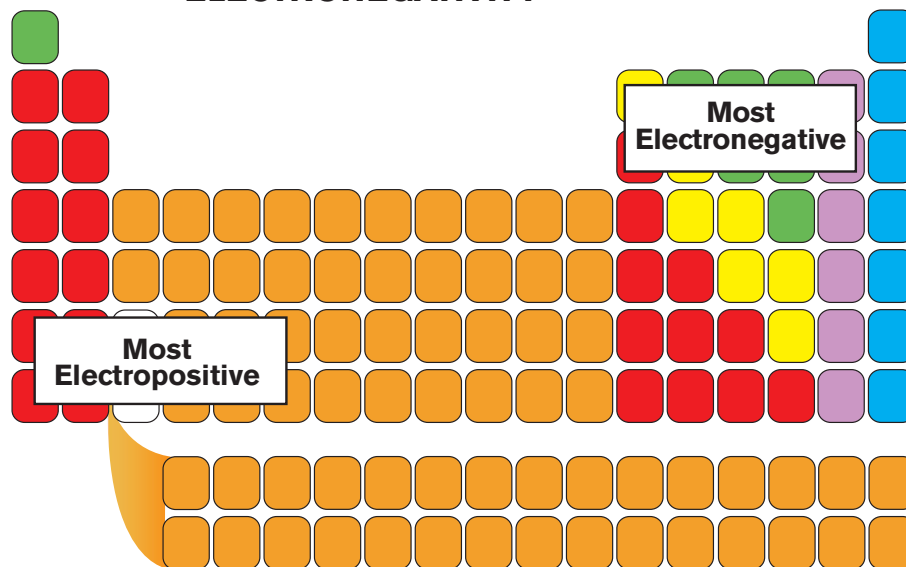
Compare sodium to its neighbor, magnesium. According to Coulomb's law, magnesium has a greater pull on the electrons. This occurs because magnesium has more protons to do the pulling. Therefore, the electronegativity of magnesium is higher than that of sodium.

- Electronegativity decreases from top to bottom down the periodic table.
- When bonding, strong electronegative elements will attract electropositive elements.

Quick Fact

Strong electronegative elements are found in the upper right of the periodic table (excluding the noble gases). Fluorine is the most electronegative element. Francium is the least electronegative. Therefore, francium is the most **electropositive** element, which means it gives up electrons most easily.

ELECTRONEGATIVITY



- **Ionization energy:** the amount of energy needed to remove the outermost electron from an atom. The most electronegative elements have the highest ionization energies.
- Other properties of elements are related to electronegativity. Going from left to right across the periodic table:
 - Atomic radii decrease.
 - Metallic properties decrease.

Quick Fact

The trend going from top to bottom down the periodic table is as follows:

Atomic radii increase.
Metallic properties increase.

TYPES OF CHEMICAL BONDS

Bonding involves an atom's valence electrons. These electrons are in the atom's outermost energy level.

EXAMPLE:

Beryllium (Be) contains four electrons. Two electrons are in energy level one (shell 1). The other two electrons are in energy level two (shell 2). The electrons in the second level are the outermost electrons for beryllium. Therefore, they are beryllium's valence electrons. These are the electrons that are involved in bonding.

Let's explore how electronegativity controls three types of bonding – ionic, covalent, and metallic.

IONIC BONDING

Ionic bonds occur when one atom gives electrons and another atom takes them. The atom that gains electrons becomes a negative ion (an anion). The atom that loses electrons becomes a cation. Therefore, the atoms have opposite charges and become attracted to one another. This force of attraction holds the atoms together.

Ionic bonds occur between atoms of elements located on opposite sides of the periodic table. Compounds held together by ionic bonds are called ionic compounds.



EXAMPLE:

When sodium (Na) and chlorine (Cl) combine to make sodium chloride (NaCl), the chlorine atoms want the extra electrons from the sodium atoms. Chlorine is on the electronegative side (the greedy side) of the periodic table. Sodium is on the electropositive side and gives away electrons to the chlorine atoms.

Step 1: $\text{Na} \rightarrow \text{Na}^+ + \text{electron}$ (production of an Na cation plus release of electron)

Step 2: $\text{electron} + \text{Cl} \rightarrow \text{Cl}^-$ (released Na electron reacts with Cl to produce a Cl anion)

Combined: $\text{Na} + \text{electron} + \text{Cl} \rightarrow \text{Na}^+ + \text{Cl}^- + \text{electron}$

Notice that the electron produced in Step 1 is used in Step 2, so it is cancelled out in the Combined reaction. (See section on [ions](#) from *The Core of Chemistry* for a review of cations and anions.)



EXAMPLE:

What would happen if magnesium (Mg) atoms were bonding with Cl atoms instead?

Step 1: $\text{Mg} \rightarrow \text{Mg}^{+2} + 2 \text{ electrons}$ (production of an Mg cation plus release of electrons)

Step 2: $2 \text{ electrons} + 2 \text{ Cl} \rightarrow 2 \text{ Cl}^-$ (released Mg electrons react with Cl to produce Cl anions)

Combined: $\text{Mg} + 2\text{-electrons} + 2 \text{ Cl} \rightarrow \text{Mg}^{+2} + 2 \text{ Cl}^- + 2\text{-electrons}$

Notice that twice as many Cl atoms are needed to take in the electrons released by the Mg. Therefore, the formula is MgCl_2 .

Quick Facts

The periodic table can be used to predict ionic compounds.

Remember: all atoms want electron configurations like the noble gases.

In the MgCl_2 example, Mg wants to be like Ne. Mg can only do this by losing two electrons. Chlorine wants to be like Ar, which only requires one electron.

Thus, two chlorine atoms are required to complete the bond.

Here's the trick:

- Count two boxes backward from Mg to get to Ne. Give the 2 to the Cl.
- Count one step forward for Cl to get to Ar. Give that 1 to the Mg.
- The result is Mg_1Cl_2 . Because we don't show the number one in formulas, we write MgCl_2 .

COVALENT BONDING

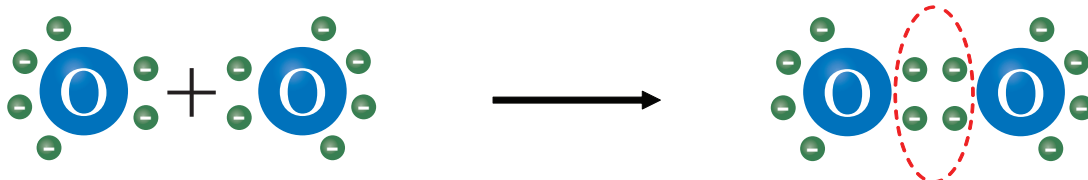
Covalent bonds occur when electrons are *shared* between two nearby atoms. Compounds formed from atoms that share electrons through a covalent bond are called covalent compounds.

- Covalent bonds create stable compounds if the sharing of electrons brings about a noble gas configuration for each atom (with 8 valence electrons).
- In a covalent bond, one atom does not actually lose an electron that is then gained by another atom. Instead, the atoms share the electrons.

EXAMPLE:

The oxygen you breathe is the gaseous molecule O_2 . Because this molecule is made of two oxygen atoms, each of the atoms equally wants to look like the nearest noble gas element, neon. The two oxygen atoms agree to share each other's electrons.

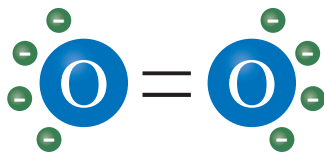
Remember that only the outermost electrons participate in the bond. Therefore, the six valence electrons in the second energy level participate in the bond. The reaction can be shown as:



The four electrons inside the dotted oval (above right) are shared. Each oxygen atom now has access to eight electrons. Therefore, both atoms appear to have the electron configuration of the noble gas, neon.

A covalent bond contains two electrons. This means there are actually two covalent bonds in an O_2 molecule (4 electrons shared, divided by 2 electrons in each bond = 2 bonds).

- Chemists often show two electrons as a line. Thus, O_2 is shown as:



- The bond that forms between the oxygen molecules to make O_2 is called a *double covalent bond*. Other chemicals may contain *single covalent bonds* or *triple covalent bonds*.
 - **Single covalent bond (single bond):** a covalent bond sharing only one pair of electrons.
 - **Double covalent bond (double bond):** a covalent bond sharing two pairs of electrons.
 - **Triple covalent bond (triple bond):** a covalent bond sharing three pairs of electrons.

Atoms do not have to be identical to form a covalent bond. They must simply be near each other and have similar electronegativities. They generally occur between nonmetal elements.

Quick Fact

Structures (such as the one at left) that show atomic centers and either lines or dots for the outer-shell electrons are called Lewis Dot Structures.

This name was given in honor of Gilbert N. Lewis for his contributions to bonding theory.

METALLIC BONDING

Metallic bonds form when the atoms in a metal contribute their electrons to a “sea” of shared electrons. This “sea” of electrons spans the entire object.

- Metallic bonds are collective by nature. A single metallic bond does not exist.
- In a metal, the outermost electrons are shared among all the atoms in the solid.
 - The creation of an electron “sea” only occurs if there is nowhere else for the electrons to go.
 - Metallic bonds tend to occur when the Coulombic forces attracting the electrons are weak in comparison to the electron energy. This allows the electrons to be easily lost by the atoms. Each atom gives up its outermost electrons, forming a “sea” of electrons.
- Elements along the left side of the periodic table often form metallic bonds.
- Metallic bonds also form in elements that have electrons with high ionization energies. These elements’ atoms do not give up electrons to other substances easily.
 - Gold, cadmium, iridium, and platinum are metals with relatively high ionization energies.
- Some metallic elements are actually quite easy to maintain in pure form. These elements include gold, copper, and silver.

Many properties of metals are a result of the high mobility of electrons in a metallic bond. They also result from the ability of those electrons to extend across the entire object.

- **Luster:** the ability of a metal to reflect light. This property gives metals a shiny appearance.
 - The large number of freely moving electrons in a metal absorb and re-emit light.
- **Electrical conductivity:** a measure of the rate at which electricity can travel through a material.
 - Metals have good electrical conductivity because their electrons can move easily throughout the metal.
- **Thermal conductivity:** the measure of the rate at which thermal energy can travel through a material.
 - Metals also have good thermal conductivity. As heat is applied to a part of the metal, the electrons become excited. When this happens, the electrons travel to the other side of the metal, carrying the energy with them. The electrons are much better at carrying the energy than the nuclei of the atoms.

Quick Fact

Silver and gold are precious metals because they are less reactive than most other metals and have a high luster. The name “precious” refers to their high economic value.

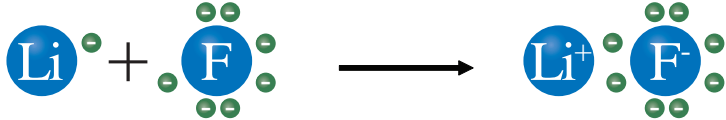
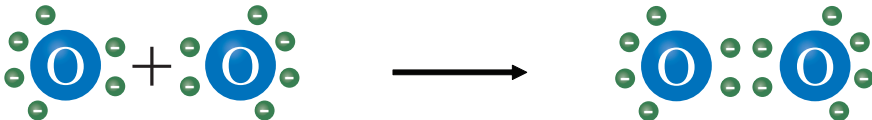
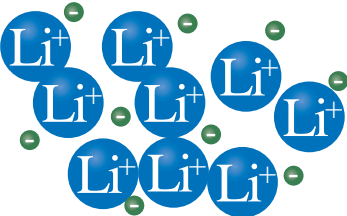
Think About It...

In the movie *A Christmas Story*, the character Flick is dared to touch his tongue to a metal flagpole in freezing temperatures. He takes the dare, and his tongue gets stuck to the metal. Why would this NOT have happened if the flagpole had been made of wood or plastic?

BONDING REVIEW

Below is a review of the three primary types of bonds – ionic, covalent, and metallic.

- Ionic bonding is essentially the result of a greedy atom combining with a giving atom.
- Covalent bonding is the result of atoms that need electrons. However, none of those atoms are greedy or giving. They share.
- Metallic bonding is the result of collectively shared electrons.

Ionic Bonding	
Covalent Bonding	
Metallic Bonding	

SECTION III: CHEMICAL REACTIONS

OBJECTIVES

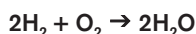
- Identify the reactants and products of a chemical reaction.
- Describe and identify examples of types of chemical reactions.
- Explain and identify reversible chemical reactions.
- Identify exothermic and endothermic reactions.
- Understand rates of chemical reactions and the effects of catalysts.

A **chemical reaction** occurs when the atoms of one or more substances are rearranged to produce one or more different substances. As a result of a chemical reaction, new substances with new properties are formed.

- **Reactants:** the starting material or materials for a chemical reaction.
- **Products:** the substance or substances produced from a chemical reaction.
- **Byproduct:** a secondary product created at the same time as the primary, desired product(s).

EXAMPLE:

The simple chemical reaction between hydrogen and oxygen is shown below:



- The hydrogen (H_2) and oxygen (O_2) are the reactants.
- Heat energy initiates the reaction.
- The resulting water (H_2O) is the product.



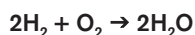
GENERAL TYPES OF CHEMICAL REACTIONS

In the chemical reactions described in this section, the letters A, B, C, and D are used to represent chemical elements and compounds.

SYNTHESIS REACTION: a chemical reaction in which two or more reactants (A and B) combine to form a product (AB).



EXAMPLE:

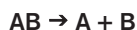


This simple reaction could power an automobile, such as the hydrogen-powered car.

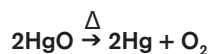
Think About It...

For this reaction to power a hydrogen car, there is plenty of oxygen in the air, but how will we get the hydrogen?

DECOMPOSITION REACTION: a chemical reaction in which a compound (AB) breaks apart into two or more products (A and B). Most decomposition reactions need an outside source of energy in order to take place.



EXAMPLE:

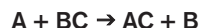


In this decomposition reaction, mercury oxide (HgO) splits into mercury metal and oxygen gas.

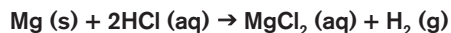
The small triangle above the arrow means the reaction needs energy to take place. (Think of the little triangle as a little flame!)

DISPLACEMENT REACTION (SINGLE REPLACEMENT REACTION):

a chemical reaction in which a reactant (A) takes the place of some part of a compound (BC). In doing so, a new compound (AC) is made, and a separate product (B) is released.



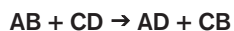
EXAMPLE:



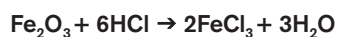
This displacement reaction happens when you combine a piece of solid (s) magnesium metal with some aqueous (aq) hydrochloric acid. When these reactants combine, they produce two products: a liquid solution called aqueous magnesium chloride and hydrogen gas (g).

DOUBLE DISPLACEMENT REACTION (DOUBLE REPLACEMENT REACTION):

a chemical reaction in which parts of two compound reactants (AB and CD) replace each other. The elements are rearranged to form two or more different compound products (AD and CB).



EXAMPLE:



In this double displacement reaction, iron oxide combines with hydrochloric acid. The reaction produces iron chloride and water.

Quick Fact

If you wanted to say what was occurring in the reaction (at left) using the English language, you would say: "Two mercury oxide molecules decompose to give us two mercury atoms plus one oxygen gas molecule."

Think About It...

How would we say the displacement reaction (at left) in English?

Quick Fact

In the world of molecules, the reaction may also look like:
 $AB + CD \rightarrow DA + BC$
because atoms do not care whether they are on the left or right side.

However, when scientists write these reactions, cations or the more electropositive element is generally written first. Therefore, the products will be written as AD and CB, if A and C represent cations.

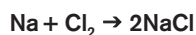
OXIDATION-REDUCTION (REDOX) REACTIONS

An **oxidation-reduction reaction** or redox reaction is a chemical reaction in which the oxidation number of the atoms change. This change is usually the result of a gain or loss of electrons (see section on [Writing Chemical Formulas](#)).

- **Oxidation:** a chemical reaction that involves an increase in oxidation number. Oxidation generally results from an atom or group of atoms losing electrons.
- **Reduction:** a chemical reaction that involves a decrease in oxidation number. Reduction generally results from an atom or group of atoms gaining electrons.

Oxidation and reduction reactions take place together. The electrons lost by one element are gained by another element. Therefore, some elements are *oxidized*, and some are *reduced*.

EXAMPLE:



The reaction between sodium and chlorine gas is a redox reaction. The complete reaction can be divided in half to illustrate the oxidation and reduction parts.

Oxidation: Sodium begins with an oxidation number of zero and ends with an oxidation number of 1⁺. It has been oxidized from a sodium atom to a sodium cation.



Reduction: The chlorine gas (Cl₂) begins with an oxidation number of zero and ends with an oxidation number of 1⁻. The chlorine atoms have been reduced to chloride anions.

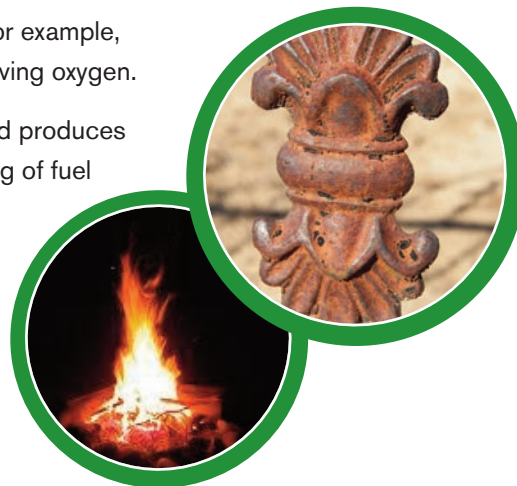


In this reaction, chlorine oxidizes the sodium atoms. Therefore, chlorine is called the *oxidizing agent*. On the other hand, sodium reduces the chlorine atoms. Therefore, sodium is called the *reducing agent*.

Redox reactions occur all around us. The combination of hydrogen and chlorine gas to form hydrochloric acid is a redox reaction. Likewise, when carbon dioxide and hydrogen gas interact to produce carbon monoxide and water, a redox reaction has occurred.

The most familiar types of redox reactions involve oxygen. For example, combustion and corrosion are types of redox reactions involving oxygen.

- **Combustion** is a redox reaction that occurs rapidly and produces energy, usually in the form of heat and light. The burning of fuel is a combustion reaction.
- **Corrosion** is a redox reaction that occurs when a metal is oxidized, usually in the presence of moist air. The rusting of iron is a corrosion process.



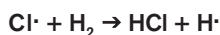
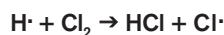
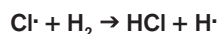
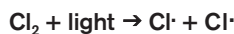
CHAIN REACTIONS

A **chain reaction** is a series of chemical reactions, in which the products of one reaction initiate further chemical reactions of the same kind. A product in the first step becomes a reactant in the second step. A product from the second step becomes the reactant for a third reaction, and so on.

- **Initiation reaction:** the chemical reaction that starts a chain reaction. The result sets up a sequence of repeating reactions.
- **Propagating reactions:** reactions that produce products that cause another reaction.
- **Termination:** the reaction or reactions that consume the substances needed to continue the reactions. At this point, the starting materials are exhausted.

Chain reactions are found in gas explosions, combustion, the formation of smog, and nuclear reactions.

EXAMPLE:



When chlorine and hydrogen interact (in the presence of light energy), a chain reaction occurs.

- **The light absorbed by a chlorine molecule breaks the molecule into special chlorine atoms, called chlorine free radicals (Cl•).**
- **The chlorine radicals then react rapidly with hydrogen molecules. The reaction produces hydrogen chloride and hydrogen free radicals (H•).**
- **The hydrogen radicals react with chlorine molecules. The reaction produces hydrogen chloride and chlorine radicals.**
- **Then, the chlorine radicals react further with hydrogen to continue the chain. This continues until some other reaction uses up the free radicals of chlorine or hydrogen.**

Quick Fact

While chain reactions often occur rapidly, some may occur slowly, such as when edible oils oxidize.

With some chain reactions, the rate of the reaction continues to increase as the number of reacting particles increases, eventually resulting in an explosion.



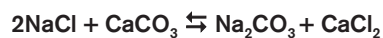
REVERSIBLE REACTIONS AND EQUILIBRIUM

Reversible reactions are reactions that can go forward (from reactants to products) or backward (from products to reactants), depending on the conditions of the experiment.

- Reversible reactions are usually represented in a chemical equation by a double arrow:



EXAMPLE:

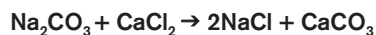


The forward reaction can be seen in saltwater lakes:



Sodium chloride (salt) in the water reacts with calcium carbonate (limestone rocks). The reaction produces sodium carbonate and calcium chloride. Calcium chloride is the salty residue seen on rocks near saltwater lakes.

The reverse reaction can be produced in a laboratory:



Sodium carbonate reacts with calcium chloride to produce sodium chloride and calcium carbonate.

- In a reversible reaction, both reactants and products may be present at the same time in a state of dynamic equilibrium.
 - **Equilibrium:** the state of a chemical reaction at which the forward and reverse reactions occur at equal rates. Therefore, the concentration of the reactants and products does not change.
 - Equilibrium describes how far a reaction goes. For instance, it describes how much product a reaction can produce (unless we manipulate it!).

Quick Fact

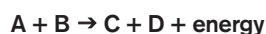
Chemists and chemical engineers often try to manipulate equilibrium. Doing this can cause a particular reaction to make more product than usual. (Therefore, they are essentially tricking the reaction).



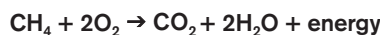
ENERGY OF CHEMICAL REACTIONS

Chemists often initiate chemical reactions to produce energy. Other times, chemists add energy to cause a reaction to take place.

- **Exothermic reactions:** chemical reactions that give off energy.
 - Exothermic reactions may occur spontaneously. They often release energy in the form of heat, light, or sound.



EXAMPLE:

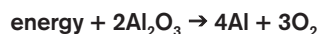


The exothermic reaction above shows how methane and oxygen produce carbon dioxide, water, and heat.

- **Endothermic reactions:** chemical reactions that require or absorb energy.



EXAMPLE:



The endothermic reaction above shows that energy is added to bauxite (aluminum oxide) to produce aluminum metal and oxygen gas.

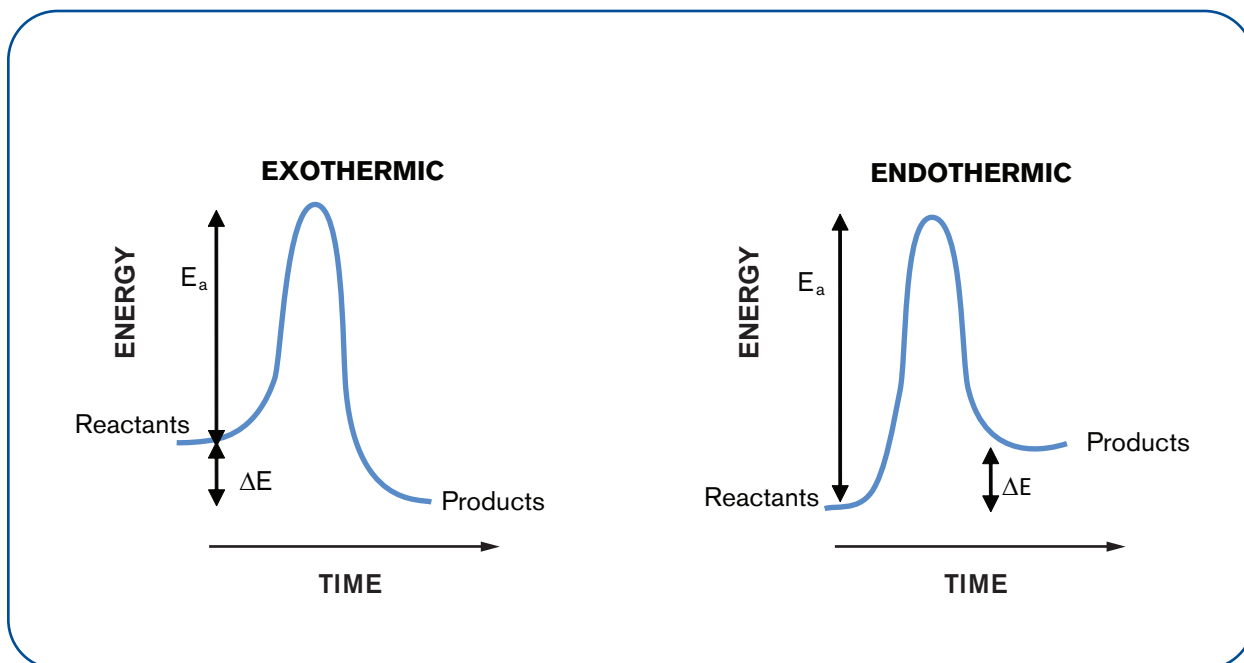
- **Energy of activation:** the amount of energy needed to cause a chemical reaction.
 - Energy of activation (activation energy) is represented by E_a .
 - As the diagrams on the following page illustrate, the activation energy is the energy required to make the reaction occur. It is the energy needed to get the reaction “over the hill.”
 - The symbol ΔE represents the change in energy. It is the difference between the starting energy of the reactants and the final energy of the products.

Quick Fact

A common example of an exothermic reaction is burning wood in a wood stove. Wood combines with the oxygen in the air to produce carbon dioxide, water, light, and heat.

Quick Fact

A common example of an endothermic reaction is the process of photosynthesis. During photosynthesis, plants use the energy from the sun to convert carbon dioxide and water into glucose and oxygen.



- In the exothermic reaction above, the reactants start at a higher energy level and end at a lower energy level. The difference, ΔE , is released from the reaction.
- In the endothermic reaction above, the products end up at a higher energy level than the reactants. This difference shows that energy had to be put into the reaction.

RATES OF CHEMICAL REACTIONS

The *rate*, or speed, of a chemical reaction is commonly affected by temperature and the concentration of the reactants and products.

- An increase in temperature usually increases the rate of the reaction.
- An increase in the concentration of the reactants increases the rate of the reaction.
- An increase in the concentration of the products decreases the rate of the reaction.

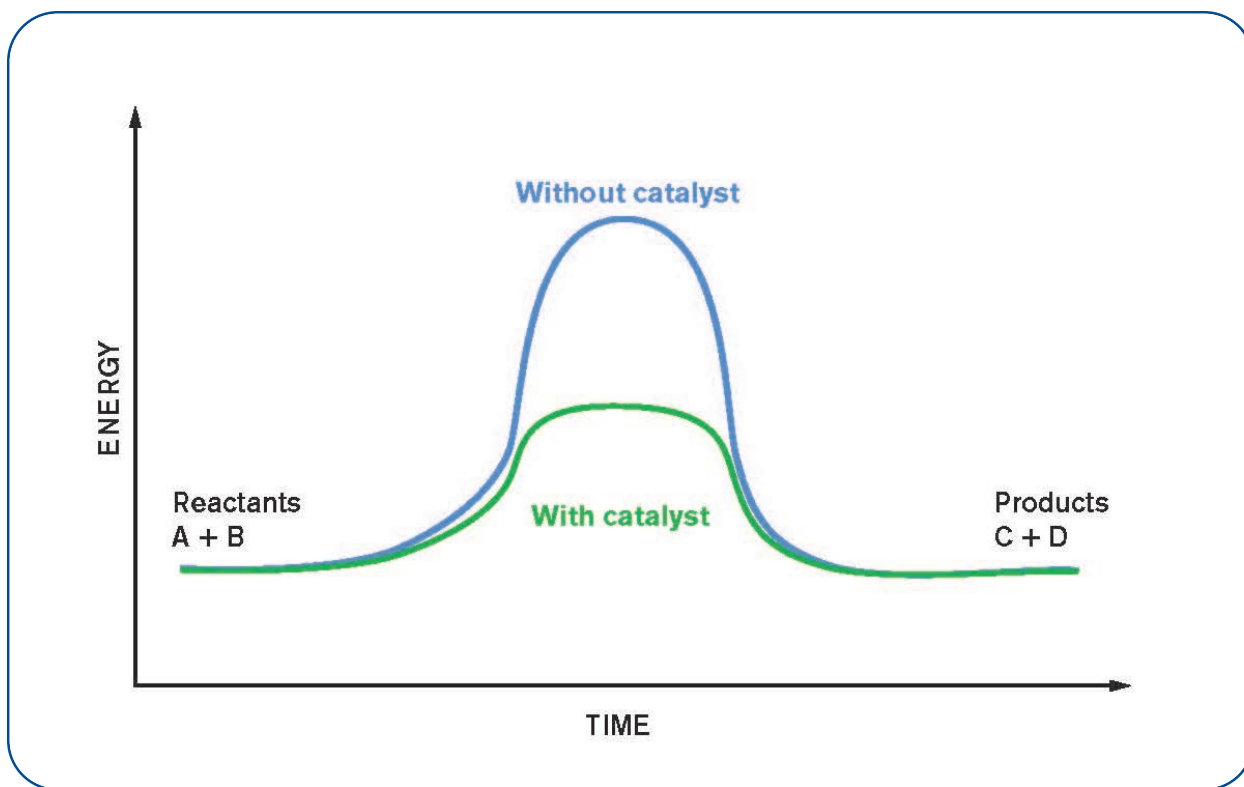
A **catalyst** is a substance that helps to change the rate of the reaction. During the reaction, the catalyst is not consumed or changed.

- The catalyst usually alters the pathway the reaction takes.
- A faster pathway usually involves a lower energy barrier. This lower barrier allows the reaction to occur at a faster rate.

Quick Fact

Humans need catalysts! Your body burns fuel (in the form of food), just like a car's engine burns fuel. However, your body doesn't require the amount of heat that a car needs to burn fuel because you have special catalysts in your body called enzymes. Enzymes allow us to burn fuel at our normal body temperature.

The diagram below illustrates the effect of a catalyst on the rate of a reaction. With a catalyst added, the “energy hill” that a reaction has to climb is much lower.



NOTES



SECTION IV: BALANCING CHEMICAL EQUATIONS

OBJECTIVES

- Describe the law of conservation of matter.
- Apply the law of conservation of matter to correctly balance equations.

When a chemical reaction occurs, it can be described by a **chemical equation**. A chemical equation uses chemical symbols and formulas to describe the reaction.

- A chemical equation shows the chemicals that react (the reactants) on the left side of the equation. The chemicals the reaction produces (the products) are on the right side.
- The chemicals are represented by their chemical symbols.
- Unlike mathematical equations, the two sides are separated by an arrow to show that the reactants form the products.

CONSERVATION OF MATTER

Law of conservation of matter (law of conservation of mass): matter cannot be created or destroyed, although it may be rearranged. According to this law, the mass of the reactants must equal the mass of the products. (Nuclear reactions are an exception.)

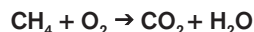
Because all matter is made of atoms, the law implies the conservation of atoms as well. This means that atoms are not lost.

What does conservation of atoms mean? The number of atoms of each element on the reactants side (left side of the arrow) must equal the number of atoms of each element on the products side (right side of the arrow). When the atoms on both sides are equal, the equation is balanced. A balanced equation demonstrates conservation of atoms.

- If a hydrogen atom goes into a reaction, it has to appear somewhere in the products of the reaction.
- Likewise, if three hydrogen atoms appear on the reactant side of a chemical equation, three must appear on the product side.

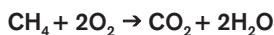
EXAMPLE:

When methane interacts with oxygen in the air, the following reaction occurs:



As written above, the reaction correctly indicates that methane and oxygen combine to form carbon dioxide and water. However, this reaction violates conservation of matter. Why? Because, there are more oxygen atoms on the right and more hydrogen atoms on the left.

The correctly balanced reaction looks like this:

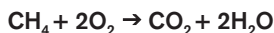


To determine the number of atoms in a chemical formula:

- Multiply the number in front of the chemical formula by the subscript number in the chemical formula.
- The number one is never written. CH_4 stands for $1\text{C}_1\text{H}_4$.
- To make sure the equation in the example above is balanced correctly:
 - Write the number of each type of atom on the reactant side.
 - Write the number of each type of atom on the product side.
 - Compare the numbers.

Quick Fact

Balancing chemical equations is like putting a puzzle together. You may not be able to tell which pieces fit where, so you may have to try a few different ways before you find a good fit. With chemical equations, you may not be able to see which numbers will work to balance the equation, so you have to experiment!

EXAMPLE:**Reactant Side of Equation**

$$\text{C: } 1 \times 1 = 1$$

$$\text{H: } 1 \times 4 = 4$$

$$\text{O: } 2 \times 2 = 4$$

Product Side of Equation

$$\text{C: } 1 \times 1 = 1$$

$$\text{H: } 2 \times 2 = 4$$

$$\text{O: } (1 \times 2) + (2 \times 1) = 4$$

HISTORY: ANTOINE LAVOISIER (1743-1794)

Antoine Lavoisier proposed the first version of the law of conservation of matter. This law states that during an ordinary chemical change, there is no noticeable increase or decrease in the quantity of matter.

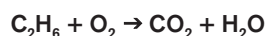
Lavoisier is known as the father of modern chemistry. He changed chemistry from a qualitative to a quantitative science.

He recognized and named oxygen. He also discovered the role oxygen plays in combustion.



GUIDELINES FOR BALANCING A CHEMICAL EQUATION

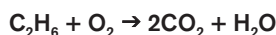
1. Write out the unbalanced equation and look to see which elements are not balanced (not equal).



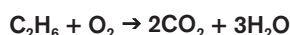
- There are 2 carbon atoms on the left side and only 1 carbon atom on the right side.
- There are 6 hydrogen atoms on the left side and only 2 hydrogen atoms on the right side.
- There are 2 oxygen atoms on the left side and 3 oxygen atoms on the right side.

2. Balance the equation. You will do this by trial and error. (Therefore, you may need to test a few different numbers before you get it right.) You will multiply the different atoms and molecules on each side by different amounts.

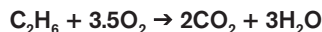
- Multiply CO_2 (on the right side) by 2. This is shown by placing a 2 in front of CO_2 . The number in front of the molecule or atom (in this case, 2) is called a **coefficient**. As mentioned before, when no coefficient (or no subscript) is written, it is assumed to be one.



- Be sure to multiply all atoms by the coefficient. Therefore, 2CO_2 means there are 2 carbon atoms and 4 oxygen atoms. **Do not change the subscripts.**
- Add a coefficient of 3 in front of H_2O on the right side, making it $3\text{H}_2\text{O}$.

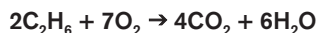


- Now, there are 2 carbon atoms on each side and 6 hydrogen atoms on each side. However, there is an uneven number of oxygen atoms (two on the left side and seven on the right side).
- Add a coefficient of 3.5 in front of O_2 on the left side, making it 3.5O_2 .



This equation is balanced, but we're not quite done.

3. A balanced equation should not contain decimals. In the equation above, the oxygen on the left is written as having a half molecule. Because there is no such thing as half an oxygen molecule, we must eliminate it from the equation. We do this by multiplying all the coefficients by two.



This equation is the properly balanced equation for the reaction.

SECTION V: CHEMICALS BY MASS

OBJECTIVES

- Define relative atomic mass.
- Identify the quantity of a mole and how it relates to the mass of atoms.

A single atom barely weighs anything. A hydrogen atom has a mass of only about 0.000 000 000 000 000 000 002 grams.

Even though atoms are very tiny, scientists still need to know the mass of an atom. In 1961, scientists started using carbon-12 as the standard atom for comparing the mass of all other atoms.

- The mass of carbon-12 was fixed at 12 atomic mass units (amu) because it has 6 protons and 6 neutrons. (Remember, protons and neutrons have almost the same mass.)
- All other masses were determined based on the weight of the carbon-12 atom.
 - **Relative atomic mass (A_r):** the weighted average mass of all an element's isotopes compared with one-twelfth the mass of one atom of carbon-12 (see section on [Isotopes](#) from *The Core of Chemistry*).

$$A_r = \frac{\text{weighted average mass of isotopes of an element}}{\frac{1}{12} \times \text{mass of 1 atom of carbon-12}}$$

or

$$A_r = \frac{\text{weighted average mass of isotopes of an element}}{1 \text{ amu}}$$

- The relative atomic mass or atomic weight of each element is listed on most periodic tables (see the Periodic Table of Atomic Weights [handout*](#)).
- For most elements, the relative atomic mass is close to a whole number. Therefore, the whole number is often used for calculations.

*The Periodic Table of Atomic Weights [handout](#) is available on the Study Materials section of CEF's website. Ask your teacher or Local Challenge Organizer for more details.

Quick Fact

A mass spectrometer is used to figure out the relative masses of atoms (as compared to carbon-12). The instrument separates ions to find out the percentage of each isotope that is present. These percentages are used to calculate the relative atomic mass of an element.

The mass spectrometer was invented by British scientist Francis Aston in 1919.

THE MOLE

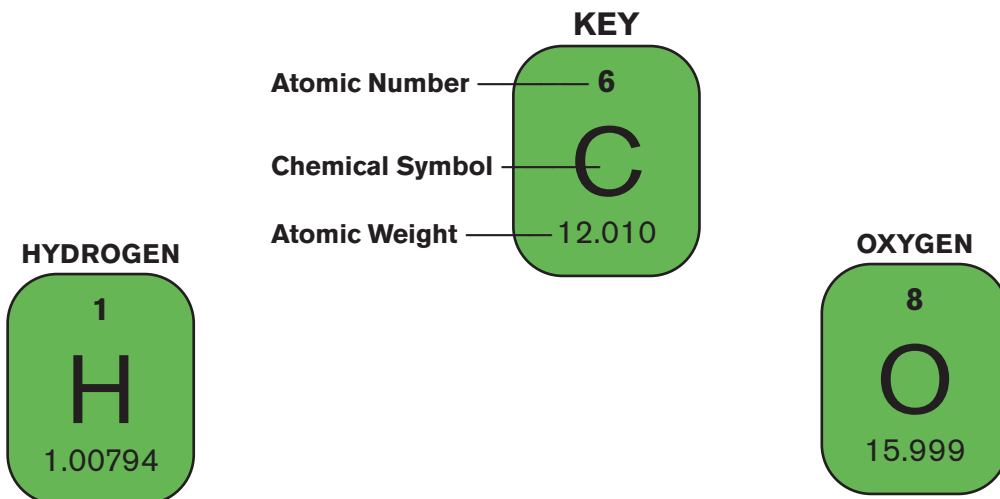
Chemists need to know more than just the mass of tiny atoms. They often want to know the number of atoms that have been used during a chemical reaction. Because atoms are so small, they cannot be counted in the same way we count larger things. So how did chemists solve this problem?

- The solution has been to weigh a large sample and then associate that weight with the number of atoms. The number chemists use is 6.0221415×10^{23} .
 - This number, called Avogadro's number, is usually estimated at 6.02×10^{23} .
 - **Mole:** an amount of substance containing 6.02×10^{23} particles.
 - Therefore, a mole of carbon is 6.02×10^{23} atoms of carbon.
- The relative atomic mass of an element in grams always contains 6.02×10^{23} atoms.

EXAMPLE:

- **A mole of hydrogen - 6.02×10^{23} atoms of hydrogen - has a mass of 1.00 grams. Likewise, 1.00 grams of hydrogen contains 6.02×10^{23} atoms.**
- **A mole of carbon - 6.02×10^{23} atoms of carbon - has a mass of 12.0 grams. Likewise, 12.0 grams of carbon contain 6.02×10^{23} atoms because carbon is 12 times heavier than hydrogen.**

Periodic tables often show an element's relative atomic mass (atomic weight) in the following format:



Quick Fact

The number written at left is a HUGE number. The number could be written out as 602,214,150,000,000,000,000,000, although, we do not actually know what those zeroes are after the 5.

Quick Fact

Chemists sometimes have special parties called Mole Day parties. Their parties begin precisely at 6:02 a.m. on October 23 (10/23).

Think About It...

If you have a mole of people, how many people are there?

Have there ever been that many people on the earth – at one point in time or over the entire history of human life on the earth?

HISTORY: AMEDEO AVOGADRO (1776-1856)

Amedeo Avogadro was an Italian chemist known for his contributions to the theory of molarity and molecular weight. He also developed Avogadro's Law.



- **Avogadro's Law** states that for equal volumes of gases at the same temperature and pressure, the gases will have equal numbers of molecules.
- As a result, the relative molecular weights of any two gases are the same as the ratio of the densities of the two gases (under the same conditions of temperature and pressure).

The concept of a mole can be extended to compounds, molecules, and other larger substances.

EXAMPLE:

Suppose you wanted to figure out how many water molecules are in a drop of water. A drop of water that you can barely see might have a diameter of about 0.1 mm. Therefore, it has a radius of 0.05 mm.

First, we can use the formula for the volume of a sphere to calculate the volume of the water droplet. The formula for the volume of a sphere is $\frac{4}{3}\pi r^3$.

$$\frac{4}{3} \pi (0.05)^3 = 0.0000005 \text{ cm}^3 \rightarrow 5 \times 10^{-7} \text{ mL}$$

In addition, we know that the density of water is 1 g/mL.

Therefore, using the density formula, we know that the mass of the droplet is 5×10^{-7} g.

Finally, using the periodic table of atomic weights, we can see that 18 grams of water contain 6.02×10^{23} molecules of water. Thus, water has a molar mass of 18 g/mol.

$$(5 \times 10^{-7} \text{ g}) / (18 \text{ g/mol}) = 2.9 \times 10^{-8} \text{ mol}$$

That's very tiny! However, if we use Avogadro's number:

$$(6.02 \times 10^{23} \text{ molecules/mol}) \times (2.9 \times 10^{-8} \text{ mol}) \\ = 1.75 \times 10^{16} \text{ molecules}$$

Thus, there are 1.75×10^{16} water molecules in the small drop of water. Written out, that's 17,500,000,000,000,000! It's no wonder scientists invented scientific notation!

Quick Fact

A *circle* is a 2-dimensional (2-D) shape in which all points on its edge are the same distance from its center.

The **diameter** of a circle is the length of a straight line that touches two points along the edge and passes through the center.

The **radius** is the distance from the center to any point on the edge.

Therefore, as the example to the left indicates, the radius of a circle is $\frac{1}{2}$ the diameter.

If you divide the circumference (distance around a circle) by its diameter, you will always get the same number. That number is pi (π) or approximately 3.142. Pi is also used to find the volume of a sphere.

A *sphere* is a perfectly round, 3-D object in which all points on its edge are the same distance from its center. Like a circle, the radius of a sphere is the distance from the center to any point on its edge.

SECTION VI: CHEMICALS BY VOLUME – SOLUTIONS

OBJECTIVES

- Describe the parts of a solution.
- Explain concentration and how to determine the concentration of a combined solution.
- Define saturation and supersaturation.
- Distinguish between polar and nonpolar solvents.
- Describe important physical properties of solutions.

Many chemicals are combined in uniform mixtures called solutions. The earth's ocean waters are massive solutions. They are made up of water, salts, and other materials.

A **solution** is a homogeneous mixture of one or more substances (the solutes) dissolved in another substance (the solvent).

- Solutions consist of elements or compounds mixed together at a very fine level. The particles of the solutes are mixed evenly with the particles of the solvent.
- **Solute:** the substance that is dissolved into the solution. The solute is mixed into the solution completely so that it seems to almost disappear. The solute is usually the substance that is present in a lesser quantity. Solute is sometimes referred to as the “active ingredient.”
- **Solvent:** the substance that does the dissolving. The solvent is usually the substance that is present in a greater quantity. Solvents are sometimes referred to as the “inert ingredient.”

EXAMPLE:

Dissolving salt into water creates a solution of salt water. In salt water, the solvent is water. The solute is salt.

Quick Fact

What happens when you mix two solutions (of the same substances) with equal volumes? The concentration of the combined solution is the average of the two starting concentrations.

For example, if you mix equal amounts of a 20% solution of salt water with a 40% solution of salt water, the combined solution would have a 30% concentration.

- **Concentration:** the amount of solute in the solution.
 - The concentration of a salt solution would tell you how much salt is in the solution.
 - To concentrate a solution, you would add more solute to a given volume. You could also remove some of the solvent. A **concentrated solution** contains a large amount of solute compared to the amount of solvent present. The solution may be referred to as a *strong* solution.
 - By contrast, to dilute a solution, you must add more solvent or reduce the amount of solute in a given volume. A **diluted solution** contains a small amount of solute compared to the amount of solvent present. The solution may be referred to as a *weak* solution.
 - Two common ways to express concentration are percent composition by mass and molarity.
 1. **Percent Composition by Mass (%):** the mass of the solute divided by the mass of the solution, multiplied by 100. (The mass of a solution is the mass of the solute plus the mass of the solvent.)

EXAMPLE:

The percent composition by mass of a 100 g solution of salt water that contains 20 g of salt is:

$$\frac{20 \text{ g salt}}{100 \text{ g solution}} \times 100 = 20\% \text{ solution}$$

2. **Molarity (M):** the number of moles of solute per liter of solution (which is not necessarily the same as the volume of solvent!).

EXAMPLE:

The molarity of a 4 liter solution containing 1 mole of salt is:

$$\frac{1 \text{ mol salt}}{4 \text{ L solution}} = 0.25 \text{ M solution}$$

- **Saturation:** the point at which no more of a solute can be dissolved into a solvent.
 - Solvents have a limit to the amount of solute they can hold at a certain temperature. If you continue to add a solute to a solution, eventually the additional solute will no longer dissolve.

EXAMPLE:

Fill a glass with a cup of water at room temperature. Add a teaspoon of salt to the water and stir until it dissolves. Continue with adding one teaspoon of salt at a time and stirring the solution until the salt dissolves. At a certain point, are you no longer able to dissolve the salt?

- This point changes significantly with different environmental factors, such as temperature, pressure, and contamination.

- **Supersaturated solution:** a solution that contains more dissolved solute than it can normally hold at a given temperature. It contains more dissolved solute than a saturated solution. A supersaturated solution is not stable. Eventually, the solute will fall out of the solution (precipitate).
 - A supersaturated solution is generally formed by cooling a saturated solution.

EXAMPLE:

Salt is added to a cup of water until the salt no longer dissolves. The solution is saturated, and excess salt rests at the bottom of the solution.

To dissolve the excess salt, you heat the solution.

At the higher temperature, water can hold more salt, so the excess salt dissolves.

However, this is a greater amount of salt than the water can normally hold at the lower temperature. Therefore, as the solution cools, it is said to be supersaturated. It is not stable, so salt crystals will easily form out of the solution.

- Crystals may form rapidly when the edge of the container is scratched or if a tiny seed crystal is added.

Solvents and solutes can be classified as either polar or nonpolar.

- **Polar solvents:** solvents made up of molecules that have an uneven distribution of electrons, creating a negative and positive side. Generally, polar solutes will only dissolve in polar solvents.
 - Water, acetone, and acetic acid are polar solvents. Salts, sugars, and ammonia are polar solutes.
- **Nonpolar solvents:** solvents made up of molecules that have an even distribution of electrons. The charges on the molecules are neutralized. Nonpolar solutes generally only dissolve in nonpolar solvents.
 - Oil, benzene, and chloroform are nonpolar solvents. Paraffin wax, hydrocarbons, and CO₂ are nonpolar solutes. (While CO₂ can dissolve in water, it is not very soluble. Think about soda pop and how it “fizzes.”)

To figure out whether a solvent will dissolve a certain solute, remember: “Like dissolves like.” This means that the more alike two molecules are, the more likely they are to mix.

Quick Fact

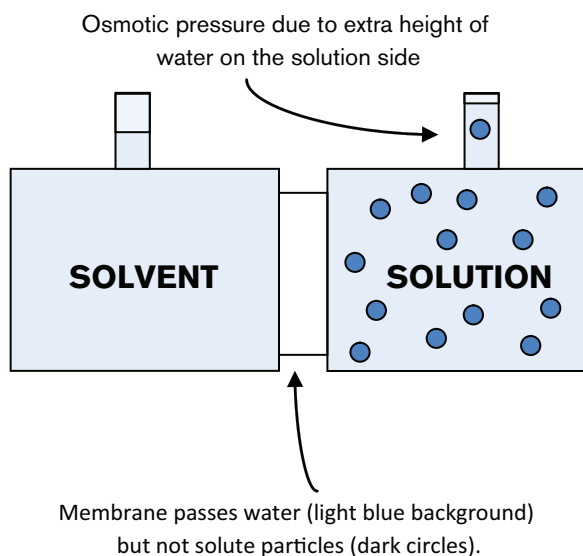
In general, as the temperature of a solvent increases, the more solute it can dissolve. However, some compounds exhibit reverse solubility. Reverse solubility means that as a solvent gets warmer, less solute can be dissolved.

For example, oxygen gas dissolves better in cold water than in hot. As a result, certain fish must live in cold water in order to “breathe.”

PHYSICAL PROPERTIES OF SOLUTIONS

Certain physical properties of solutions make them very useful to chemists. Some of these important physical properties are described in the table below:

Physical Property of Solutions	Description of Property
Freezing point depression:	Solutions freeze at a lower temperature than the pure solvent. Antifreeze for automobiles freezes at a lower temperature than the pure solvent, water. (Antifreeze is a solution made of water and ethylene glycol.)
Boiling point elevation:	Solutions boil at a higher temperature than the pure solvent. Adding salt to water causes the solution to boil at a higher temperature. This decreases the amount of time it takes to cook food. This property of solutions is important at high elevations. Without adding salt, the water may boil at such a low temperature that the food might not be cooked properly.
Osmotic pressure:	Certain membranes permit the passage of a solvent, but not the solute. When such membranes are used to separate a solution and a solvent, more solvent flows into the solution. Eventually, the solution's weight builds up enough pressure to stop the flow.





SECTION VII: ACIDS, BASES, AND pH

OBJECTIVES

- Explain pH and describe substances as acidic or basic based on the pH scale.
- Define and identify common acids and bases.
- Identify and describe common indicators.

THE pH SCALE

The **pH scale** is used to measure the acidity of a solution.

- Acids release hydrogen ions [H^+] when dissolved in water. Thus, the acid content of a solution is based on the concentration of hydrogen ions in the solution.
- The pH scale is the tool used to indicate the concentration of hydrogen ions in a solution.
- Usually, substances range from 0 to 14 on the pH scale.
 - The smaller the number on the pH scale, the more acidic the substance is. Thus, a substance with a pH of 1 is very acidic. It is a very strong acid.
 - The more basic a substance is, the higher its number on the pH scale. Thus, a substance with a pH of 13 is a very strong base.
 - Pure (distilled) water has a neutral pH of 7.0. A neutral substance is neither acidic nor basic.
 - Some very strong acids may have a pH less than 0. For example, concentrated hydrochloric acid (HCl) may have a pH of zero or slightly less than zero.

Quick Fact

Small changes on the pH scale actually mean large changes in acidity. A change in just one unit (from pH 6.0 to pH 5.0) indicates that its acidity has increased by a factor of 10.

For example, if the pH of a substance decreases by 3 (from 6.0 to 3.0), the acidity has increased by 1,000.



The table below lists some common acids and bases on the pH scale.

Substance	Approximate pH	Approximate pH Indicator Paper Color
Sulfuric acid, battery acid	0.8-1.5	Red
Stomach acid	1.0-2.0	Red
Lemon juice, cola	2.3-2.5	Orange
Vinegar	2.9	Orange
Apple juice, orange juice	3.3-3.8	Yellow
Coffee	5.0-5.5	Yellow
Milk	6.5	Light Green
Pure water	7.0	Green
Human blood	7.4	Dark Green
Sea water	8.0	Teal
Baking soda solution	8.5-9.0	Dark Teal
Milk of magnesia	10.5	Dark Blue
Household ammonia	11.5-12.0	Dark Blue
Bleach	12.5	Dark Blue
Liquid drain cleaner	13.5-14.0	Black

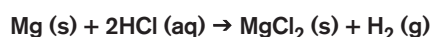
ACIDS

Acids are solutions (or chemical compounds dissolved in water) that have an excess of hydrogen ions (H^+).

- Acids are able to give up H^+ ions to bases.
- Acids can conduct electricity and are corrosive in nature. They have the ability to dissolve some metals.
 - When an acid reacts with a metal, it produces a metal salt and hydrogen.

EXAMPLE:

When magnesium comes in contact with hydrochloric acid, the acid reacts with the metal. The reaction produces magnesium chloride (a salt) and hydrogen gas.



Quick Fact

The word “acid” comes from the Latin term “acidus,” which means sour. Acids generally have a sour taste.

Remember, you should never taste a substance to determine what it is!

Quick Fact

Clean rain usually has a pH of 5.6, which is slightly acidic because of the carbon dioxide that is naturally present in the atmosphere. Rain measuring less than 5 on the pH scale is abnormally acidic and therefore, called *acid rain*.

BASES

Bases are solutions (or chemical compounds dissolved in water) that have an excess of hydroxide ions (OH^-). They will accept H^+ ions from acids.

- Likewise, bases are able to donate OH^- ions to acids.
- Bases feel slippery to the touch and are often used to make soaps.
 - However, strong bases, such as drain cleaner, can be dangerous to your skin.
- Although the term “alkali” is often used as a synonym for base, they are not the same thing.
 - Alkalis are basic, ionic salts of an alkali metal or an alkaline earth metal. Therefore, *all alkalis are bases, but not all bases are alkalis.*

Quick Fact

Bases typically have a bitter taste and, like acids, can conduct electricity.

EXAMPLE:

Calcium carbonate and soda lye are bases that are also alkali salts.

Ammonia is a base but *not* an alkali.

STRENGTH OF ACIDS AND BASES

Acids and bases may be strong or weak depending on how well an acid or base produces ions in water.

- A strong acid produces many hydrogen ions. A weak acid produces fewer hydrogen ions. As a result, indicator paper and litmus paper register slightly different colors depending on the strength of the acid (see section on [Indicators](#)).
- The strength of an acid may be affected by the size of the anion (see section on [Ions](#) from *The Core of Chemistry*) produced when the hydrogen is released into water.
 - Larger anions are more stable. They are more easily separated from the hydrogen ion.
 - In general, strong acids include hydrochloric acid, nitric acid, sulfuric acid, and hydrobromic acid.
- Strong bases act in a manner similar to strong acids, producing hydroxide ions instead of hydrogen ions.
 - In general, strong bases include sodium hydroxide, potassium hydroxide, and lithium hydroxide.



INDICATORS

Indicators are substances that change color at a specific pH. They provide a way to determine the acidity of a solution. Some common indicators are:

- **Litmus paper:** an indicator that turns red in an acidic solution or blue in a basic solution.
- **Phenolphthalein solution:** an indicator that changes from clear to pink for a pH greater than 9.
- **Bromothymol Blue (BTB):** an indicator that turns yellow in acidic solutions and blue in basic solutions.

Quick Fact

Certain foods and flowers act as indicators. Cherries and beets appear red in acidic solutions but turn blue or purple in basic solutions. The flowers of hydrangea plants are blue in acidic soil but pink or white in basic soil.



NOTES



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