You Be The Chemist® Challenge

PASSPORT TO SCIENCE EXPLORATION
CHEMISTRY CONNECTIONS

CHEMICAL EDUCATIONAL FOUNDATION®

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

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

The information in this *Passport to Science Exploration* will expose you to the fascinating world of chemistry as it applies to your everyday life. To prepare for the Challenge, read and review the information in these study materials to further your understanding of chemistry. Challenge competition questions are based on an overall understanding of chemistry and general science concepts and the relationships among concepts. Challenge questions are largely based on the 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 portions that correspond to different levels of the Challenge: *The Core of Chemistry* (local), *Chemistry Connections* (state), and *Chemistry Concepts in Action* (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* for State Challenge competitions[1], 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.

We encourage you to explore *all* the study materials provided on CEF's website at 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!

Below are some tips to help you get the most out of the information found in these study materials.

1. Read over the *Table of Contents* first to see what concepts are covered.
2. Review the *Objectives* list provided at the beginning of each section to become familiar with the topics you will learn. You shold go back and review the *Objectives* once you have finished reading a section. Can you complete the tasks listed in the *Objectives*? If not, review the section again.
3. Know the definitions of the **bolded terms.** Go back and review these after reading a section.
4. Use the diagrams, pictures, and illustrations to gain a better understanding of the concepts.
5. Read through the "Quick Facts" in each section. Quick Facts are not necessary for understanding the major concepts but they provide details and applications to help you understand the material even better.
6. Read the information in the History boxes. These boxes provide a variety of background knowledge about chemistry concepts and help to connect science of the past to science today.

[1]The 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 or State Challenge Organizer in the event of such changes.

7. Use the information in the Element boxes to learn more about specific elements.

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 them 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. If you find that you still do not understand something, ask your science teacher, or pull out a science textbook and look in the index for the information. You can also conduct a search on the Internet. (Be sure to find a reliable source.)

TABLE OF CONTENTS
CHEMISTRY CONNECTIONS

I. CHEMICAL FORMULAS

II. FORCES OF ATTRACTION

III. CHEMICAL REACTIONS

IV. BALANCING CHEMICAL EQUATIONS

V. ACIDS, BASES, AND pH

VI. RADIOACTIVITY & NUCLEAR REACTIONS

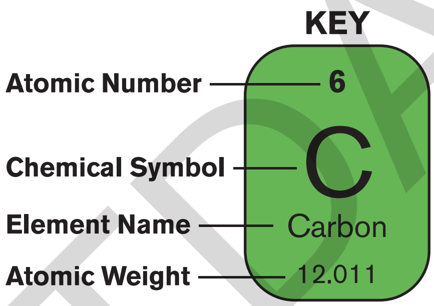
SECTION I:
CHEMICAL FORMULAS

OBJECTIVES

- Identify the chemical symbols and Lewis structures for different elements.
- Identify and write chemical formulas using chemical symbols.
- Recognize common chemical compounds and their formulas.
- Use the guidelines of chemical nomenclature to name chemical compounds.

CHEMICAL SYMBOLS

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 either 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.



LEWIS STRUCTURES

Scientists have many different ways to represent an atom of an element. One way is to write the element's ground state electron configuration (see the subsection on Electron Configuration from *The Core of Chemistry*). Another way is to use Lewis structures. Lewis structures (also known as electron dot structures) contain the element's chemical symbol and dots that represent the high-energy outermost (or valence) electrons.

Valence electrons are located the farthest from the nucleus of an atom (see Periodic Trends). Valence electrons are often involved when reactions occur, so looking at a Lewis structure and knowing how many valence electrons an element has can help determine how it will interact with other elements.

Quick Fact
The ground state electron configuration of an atom shows the lowest energy state of the atom.

To draw the Lewis structure of an element, you first have to determine how many valence electrons an atom of that element has. You can do this by using the periodic table and locating the group in which the element is found. The electron dots are then placed around the four sides of the chemical symbol as follows:

- First, place a single dot for each valence electron around each side of the chemical symbol.
- After you have placed one dot on each side, the dots can be paired with another electron dot until all valance electrons are shown in the structure.

Atom of Element	Lewis Structure
Lithium	Li•
Boron	•Ḃ•
Carbon	•Ċ• (4 dots)
Fluorine	:F: (7 dots)
Chlorine	:Cl: (7 dots)
Bromine	:Br: (7 dots)

Most main group elements will have up to eight valence electrons, but transition metals do not follow this rule.

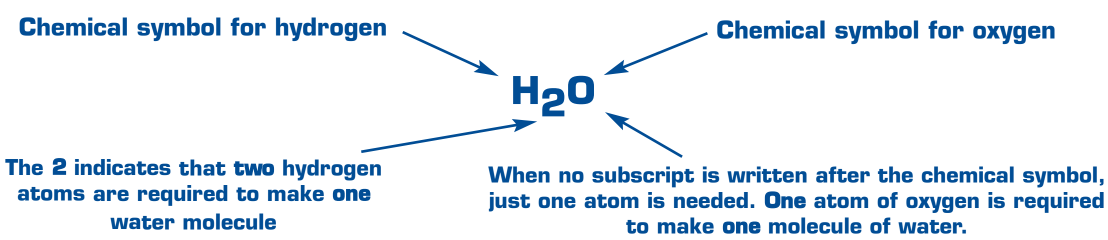
If you draw the Lewis structures for elements in the same group, you can see that they look the same. This can help you understand why elements of the same group tend to react similarly. Look at the Lewis structures for fluorine, chlorine, and bromine in the table above. All of the halogens (group 17) need just one more electron to have a full octet and be stable, so they'll react with elements that will complete their octets. (See the subsection on Types of Chemical Bonds.)

WRITING CHEMICAL FORMULAS

A **chemical formula** of a compound shows the type and number of atoms of an element that are part of that compound. Remember, a compound is a pure substance made up of two or more elements that are chemically combined.

EXAMPLE:

Water is a compound that contains the elements hydrogen and oxygen. Two hydrogen atoms and one oxygen atom are needed to form one water molecule, so the chemical formula is H_2O.



Chemical formulas use **subscripts** to indicate how many atoms of each element there are in a given compound. Subscripts are the numbers located at the lower right of a chemical symbol. If you look at the subscripts after hydrogen and oxygen in the previous example, you can see how many atoms of each element are needed to make a molecule of water: 2 hydrogen atoms and 1 oxygen atom.

Chemical formulas also show the total number of atoms that are in one molecule of a given compound. Water has a total of three atoms per molecule – two hydrogen atoms and one oxygen atom.

Oxidation numbers and charges are also important when writing chemical formulas. Oxidation numbers are used to show how many electrons an atom gains, loses, or shares when it chemically combines with another element. It is important to remember the following when thinking about charges and oxidation numbers of elements:

- Typically when the outermost energy level/sublevel of an atom contains eight valence electrons (like the noble gases), the atom is very stable (see the subsection on Periodic Trends).
- Elements want to be stable and have eight valence electrons. Atoms of elements cannot simply get rid of or gain electrons on their own, so they have to *gain, lose, or share electrons by forming chemical bonds with other atoms*.
- Remember that electrons carry a negative charge, so when a neutral atom undergoes a reaction and gains electrons, it end up with a negative oxidation number. When a neutral atom loses electrons, it will have a positive oxidation number. The noble gases in group 18 (inert gases), which have a full octet, have an oxidation number of zero.

When writing chemical formulas, the element with the positive oxidation number is written first. Since metals tend to give up electrons, they usually have positive oxidation numbers and are written first. Nonmetals typically have negative oxidation numbers.

Consider writing the formula for sodium chloride (NaCl). From the periodic table, you can see that sodium needs to give away an electron to complete its octet. When sodium bonds with chlorine, it will give away an electron and result with an oxidation number of +1. Chlorine will take the additional electron to complete its octet and obtain an 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 (see the subsection on Naming Chemical Compounds).

This change results from gaining an electron. It reflects the change (decrease) 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 an element that has multiple possible oxidation numbers, scientists use Roman numerals to show the oxidation number that it has in a given compound. Iron (II) has oxidation number +2. Iron (III) has one +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. 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 $H_2CO_3\text{ (aq)} \rightarrow H_2O\text{ (l)} + CO_2\text{ (g)}$, carbonic acid, H_2CO_3, is in 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 it may also contain other chemical substances, such as magnesium carbonate. Many brands of table salt also contain additives, such as iodine, for health reasons.

NAMING CHEMICAL COMPOUNDS

Chemical nomenclature is the system used for naming chemical substances. There are millions of identified chemical substances, so naming them all would be difficult without a set of systematic rules. Some compounds have been recognized for a long time and have common names (like water) but most compounds do not have common names. Instead, they have standard names based on the naming rules established by the International Union of Pure and Applied Chemistry (IUPAC). The rules for naming compounds depend on the type of compound, so similar compounds, like acids or ionic compounds, have similar names.

Quick Fact

There are exceptions for naming compounds. Using the guidelines for naming compounds, we might expect the chemical name of water to be dihydrogen monoxide. However, that name is not used by scientists. The IUPAC name for water is oxidane. Yet, water is most commonly called "water" even by scientists.

NAMING IONS AND IONIC COMPOUNDS

An **ion** is an atom or molecule that has lost or gained one or more of its outer electrons. Ions may have a positive charge (a cation) or negative charge (an anion). Ions of a single element are called **monatomic ions**.

- Monatomic *cations* formed from metal atoms keep the same name as the element.

Symbol/Formula	IUPAC Name
K^+	Potassium ion
Na^+	Sodium ion
Al^{3+}	Aluminum ion
H^+	Hydrogen ion

- If a metal can form different *cations*, the positive charge is shown by placing Roman numerals in parentheses after the metal name. Many of the transition metals can have different oxidation numbers and still be stable. Take a look at the ions of iron in the table below. Both cations are stable and are likely to form.

Symbol/Formula	IUPAC Name
Fe^{2+}	Iron (II) ion
Fe^{3+}	Iron (III) ion

If you are unsure if a metal forms more than one cation, you can use Roman numerals to indicate the charge even if it is unneccesary. For example, although Al^{3+} is a more common ion, the oxidation states +1 and +2 are also possible. We can write aluminum (III) ion to refer to the ion with a +3 oxidation.

To distinguish anions from cations, anions have different endings to their names.

- Monatomic *anions* are named by replacing the ending of the element name with "–ide."

Symbol/Formula	IUPAC Name
H^-	Hydride ion*
Cl^-	Chloride ion
N^{3-}	Nitride ion

*As noted in *The Core of Chemistry*, H^- is a hydrogen anion. These anions are officially called hydride ions, as listed above.

Ionic compounds are formed from ions bonded together by electrical forces (see the subsection on Intramolecular Forces). Ions (ionic compounds) of more than one element are **polyatomic ions** or molecular ions.

- Polyatomic *cations* formed from nonmetal atoms have names that end in "–ium."

Symbol/Formula	IUPAC Name
NH_4^+	Ammonium ion
H_3O^+	Hydronium ion

- Polyatomic *anions* may be written in different ways depending on the number of atoms and the elements that combine. For example, some are named in the same way as monatomic ions with the "–ide" ending (suffix).

Symbol/Formula	IUPAC Name
OH^-	Hydroxide ion
CN^-	Cyanide ion

- Polyatomic *anions* that contain oxygen have names that end in "–ate" or "–ite." These anions are known as **oxyanions**. The suffix "–ate" is used for the most common oxyanion of an element. The suffix "–ite" is typically used for an oxyanion with the same charge but with one less oxygen atom.

Symbol/Formula	IUPAC Name
SO_4^{2-}	Sulfate ion
SO_3^{2-}	Sulfite ion

- An ion that contains hydrogen and an oxyanion is named by adding the word hydrogen or dihydrogen as a prefix (at the beginning).

Symbol/Formula	IUPAC Name
HCO_3^-	Hydrogen carbonate ion
$H_2PO_4^-$	Dihydrogen phosphate ion

As you may have noticed, ionic compounds are named by writing the cation name first and then the anion name.

NAMING ACIDS

Acids are chemical compounds that give up hydrogen cations (H^+) when they are dissolved in water (see the Acids, Bases, and pH section). Therefore, acid compounds contain hydrogen.

Some acid compounds have only two elements - hydrogen and another element. They are called *binary acids*. These acid compounds do not contain oxygen.

- To name *binary acids*, you can use the following steps:

  1. Start with the prefix "hydro–."
  2. Then, add the name of the monatomic anion (of the other element).
  3. Next, change the "–ide" ending to "–ic."
  4. Finally, add the word "acid."

EXAMPLE:

For the acid HF, you begin with the prefix hydro. Then, add on the name of the monatomic anion of fluorine, which is fluoride. When you put the first two parts together, you get hydrofloride. Next, change the "–ide" ending to "–ic." Now you have hydrofluoric. Finally, add the word "acid." The name of the compound HF is hydrofluoric acid.

Symbol/Formula	IUPAC Name
HCl	Hydrochloric acid
H_2S	Hydrosulfuric acid

Other acid compounds contain oxygen. These acids are called *oxyacids*. There are two main methods for naming acid compounds that contain oxygen.

- To name *oxyacids that contain an anion ending in "–ate,"* you can use the following steps:

  1. Start with the name of the anion.
  2. Next, change the "–ate" ending to "–ic."
  3. Finally, add the word "acid."

EXAMPLE:

For the acid $HClO_3$, you begin with the name of the anion. The anion ClO_3 is a chlorate ion. Next, change the "–ate" ending to "–ic." Now, you have chloric. Finally, add the word "acid." The name of the compound $HClO_3$ is chloric acid.

Symbol/Formula	IUPAC Name
HNO_3	Nitric acid
H_2SO_4	Sulfuric acid

- To name *oxyacids that contain an anion ending in "–ite,"* you can use the following steps:
  1. Start with the name of the anion.
  2. Next, change the "–ite" ending to "–ous."
  3. Finally, add the word "acid."

EXAMPLE:

For the acid compound $HClO_2$, you begin with the name of the anion. The anion ClO_2 is a chlorite ion. Next, change the "–ite" ending to "–ous." Now, you have chlorous. Finally, add the word "acid." The name of the compound $HClO_2$ is chlorous acid.

Symbol/Formula	IUPAC Name
HNO_2	Nitrous acid
H_2SO_3	Sulfurous acid

NAMING MOLECULAR COMPOUNDS (INORGANIC)

Molecular compounds are often formed by combining two or more nonmetal elements. They are typically held together by covalent bonds (see the section on Types of Chemical Bonds).

- To name molecular compounds from the molecular formulas, you can use the following steps:
  1. Write the name of the element that is the most electropositive. This is usually the element that is further to the left in the periodic table. If both elements are in the same group, then the element with the higher atomic number is usually written first.
  2. Write the name of the second element.
  3. Change the ending of the second element to "–ide."
  4. Finally, use Greek prefixes (below) to represent the number of atoms of each element in the compound.

Think About It...

After reading the section on Periodic Trends, what can you conclude about the electronegativity of the atom element that is written first? Should it have a higher or lower electronegativity than the second element?

Prefixes	
1: mono-	6: hexa-
2: di-	7: hepta-
3: tri-	8: octa-
4: tetra-	9: nona-
5: penta-	10: deca-

Quick Fact

The Greek prefix mono– is not used when there is only one atom of the *first* element in a molecular compound. That is why CO is called carbon monoxide rather than monocarbon monoxide!

EXAMPLE:

For the molecular compound CO, carbon has a positive oxidation number (and is found farther to the left on the periodic table). Therefore, carbon is written first. Oxygen is written next, and its ending is changed to "–ide." Now you have oxide. Finally, add the Greek prefix "mono–" to oxygen. (Mono means one. Monoxide means that there is only one oxygen atom.) Thus, the name of the compound is carbon monoxide.

Likewise, the name of the molecular compound CO_2 is carbon dioxide.

Symbol/Formula	IUPAC Name
Cl_2O	Dichlorine monoxide
NF_3	Nitrogen trifluoride

Quick Fact

Be careful not to confuse molecules with ions. Some molecules have almost the same formula as certain ions. For example, the molecular formula for sulfur trioxide is SO_3. The formula for a sulfite ion is SO_3^{2-}. The formula used to represent an ionic compound is called a formula unit. The formula used to represent a molecular compound is called a molecular formula.

NOTES

SECTION II:
FORCES OF ATTRACTION

OBJECTIVES

- Explain the relationship of Coulomb's law and electronegativity to chemical bonding.
- Describe the periodic trends for electronegativity, ionization energy, and atomic radii.
- Identify the three primary types of bonds.
- Use Lewis structures to illustrate bonding.
- Explain the forces of gravity and magnetism.

Although forces are most often discussed in relationship to work and pressure, a force is any kind of push or pull on an object. Work and pressure are types of forces that require objects to touch, so they are known as *contact forces*. Friction, air resistance, and tension are also examples of contact forces. There are other forces that can act on objects at a distance; these are known as *non-contact forces*. For example, the most familiar unseen, non-contact forces are gravity and magnetism.

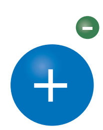
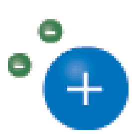
Although the forces mentioned above are easily seen and experienced, there are other types of forces that play an important role inside chemical substances, including intermolecular and intramolecular forces. An **intermolecular force** is a force acting between two or more molecules. An **intramolecular force** is the force of attraction *within* atoms, ions, or molecules.

INTRAMOLECULAR FORCES

Intramolecular forces of attraction hold together atoms and molecules. These forces are described by **Coulomb's law** which states that:

- 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 in at atom push each other apart if they are the same (both positive or both negative).

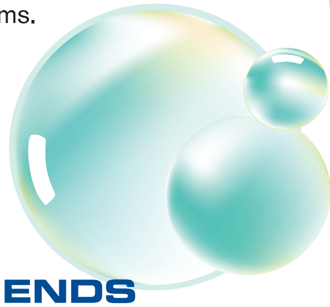
For example, two electrons will repel each other because they both have negative charges, but an electron and a proton will attract each other because one is negative and one is positive.

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

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

These forces also apply to ions and ionic compounds. For example, an ammonium ion (NH_4^+) will be attracted to a chlorine anion (Cl^-). A bromine anion (Br^-) will be attracted to a potassium cation (K^+).

Coulomb's law and other laws of chemistry combine explain the structure of atoms.



Quick Fact

Although Coulomb's law plays an important role in atoms, were it not for the laws of quantum mechanics, which we will not discuss here, protons and electrons would simply crash into each other because they are attracted. Atoms as we know them could not exist. Is it quantum mechanical laws that keeps electrons in an atom at a certain distance from the nucleus, and it is the balance of Coulombic forces and quantum mechanics that keeps atoms and molecules stable.

PERIODIC TRENDS

As mentioned previously, **valence electrons** are the electrons in the outermost energy level of an atom (also taking into account sublevels). They are represented as the dots that surround the chemical symbol in a Lewis structure.

- An atom of a main group element can typically hold eight valence electrons. The exceptions are hydrogen and helium.
- Atoms that have fewer than eight valence electrons tend to form bonds with other atoms. They will give, take, or share electrons to achieve a full outermost energy level (with eight eletrons), which will make them stable.
- Although some energy levels can hold more than eight electrons, main group elements have a maximum of eight valence electrons that participate in chemical bonding. For example, energy level three (which has sublevels 3s, 3p, and 3d) can hold a maximum of 18 electrons (3s=2 electrons, 3p=6 electrons, 3d=10 electrons). However, the element chlorine ($1s^22s^23p^6\mathbf{3s^23p^5}$) will only use 7 valence electrons in a chemical bond (see the section on Electron Configuration from *The Core of Chemistry*).

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 valence electrons.

In addition, the energy level of an element's valence electrons relates to the period on the periodic table in which that element is found. For example, potassium is in the fourth period, so this means that its valence electrons are on the fourth energy level. For the main group elements, the valence electrons are on the energy level that relates to the period in which they are found.

The number of valence electrons for atoms of elements in groups 13–18 can be determined in a similar way. These atoms have ten fewer valence electrons than their elemental group number. For example, atoms of elements in group 13 have three valence electrons. Atoms of elements in group 18 (noble gases) have eight valence electrons which explains why they are relatively unreactive.

The process is not so simple for the transition metals in groups 3–12. The number of valence electrons for these elements' atoms can vary.

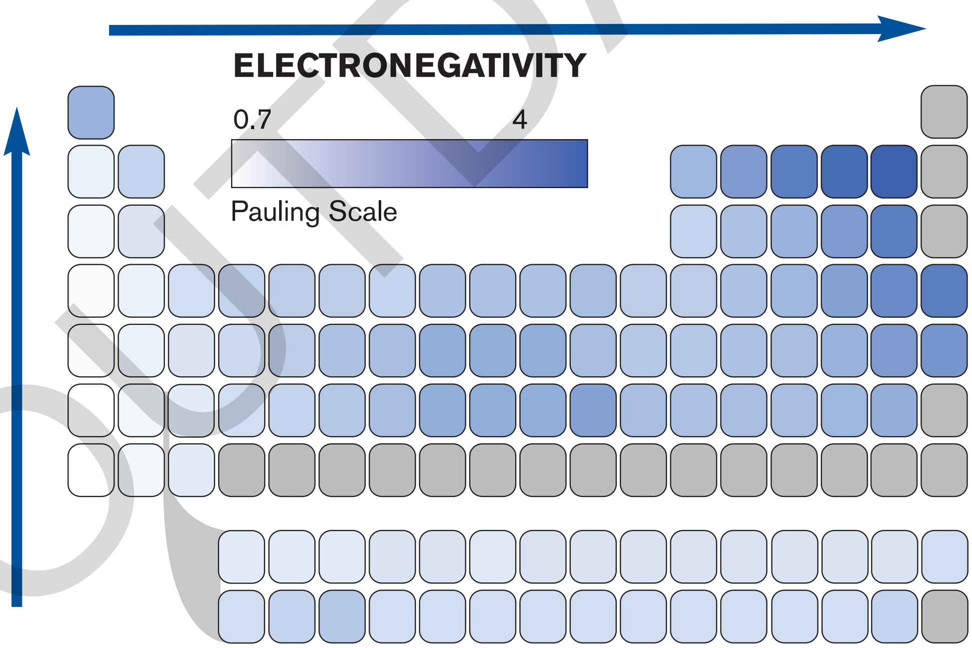
ELECTRONEGATIVITY

Electronegativity is a chemical property that describes a measure of how strongly the nucleus of an atom attracts electrons when bonding with other atoms. Electronegativity essentially measures how strong of a pull an atom has on electrons and how much it will compete for electrons during reactions. When atoms bond, the most electronegative atom will attract the electrons.

- **Within a period, electronegativity *increases* from left to right.** If you are on the same period of the periodic table, moving to the right means that more protons, neutrons and electrons will be added to the same space (because it's within the same shell). You could think of the proton-electron attractions as small magnets attracting each other. The more proton/electron pairs you have, the more the outer shell will be drawn towards the center of the atom. As the outer shell is drawn in and the valence electrons get closer to the nucleus, the atom is able to hold on to its own and other electrons more tightly.

  The nucleus of smaller atoms can hold on to the electrons better than a larger atom can, so when a smaller atoms interacts with other atoms it will be more capable of adding electrons to its shell to complete its octet. The smaller atom's ability to attract and hold on to additional electrons signifies that it is more electronegative than an atom that is larger in size. (See the subsection on Atomic Radii.)

- **Within a group, electronegativity *decreases* from top to bottom.** This can also be explained by looking at the size of the atoms. As you go down a group, you add more shells to the atoms, which increases the size. As the atoms get bigger (as you go down a group), there is more space and more electrons between the nucleus and the outer shell, so the nucleus can't attract other (or its own) electrons very well compared to a smaller atom. The valence electrons are also shielded from the nucleus by the additional shells that are added moving down a group.



*Electronegativity is measured on the Pauling scale, with 0.7 being the least and 4.0 being the most electronegative.
**The electronegativity is unknown for the elements on the table shown in gray.
***The arrows indicate the *general trend* of electronegativity. It increases as you move to the right within a period and as you move from the bottom to the top within a group.

Notice that 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, which means it gives up electrons most easily. Francium is also said to be the most "electropositive" element.

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. Since the outer shell is full, neon does not need to gain or lose electrons - it is already stable!

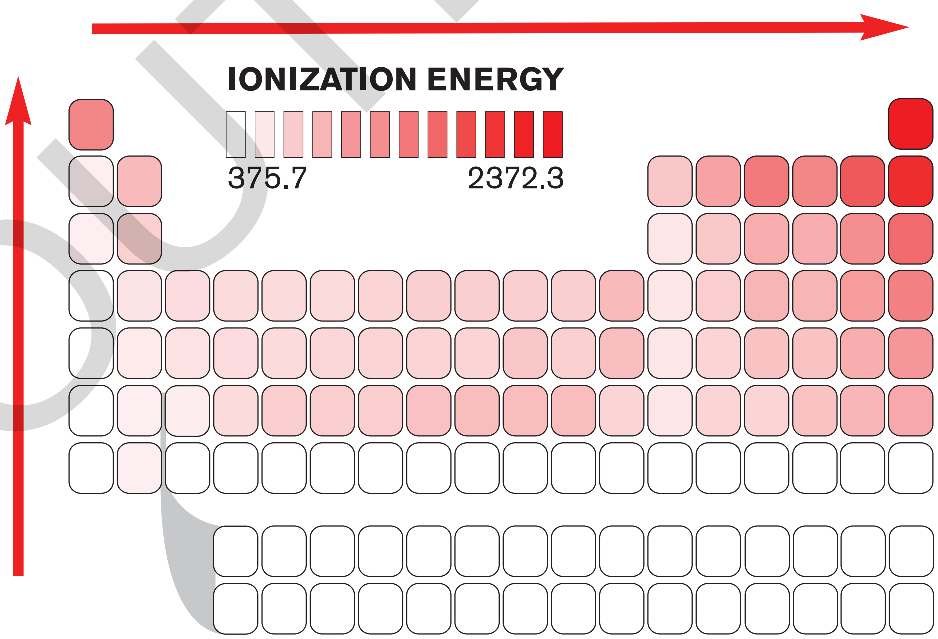
EXAMPLE:

Compare sodium (Na) to its neighbor magnesium (Mg). 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.

IONIZATION ENERGY

Ionization energy is the amount of energy needed to remove the outermost electrons from a neutral atom. It is the energy needed to overcome the attraction of the negative electron to the positive nucleus. Think of it as an indicator of how strongly an atom holds on to its valence electrons. High ionization energy means the atom has a strong hold on its valence electrons, so it will take more energy to get the electrons away from the atom/nucleus. Low ionization energy means that the atom can easily lose its valence electrons, compared to other atoms.

- **Within periods, ionization energy tends to *increase* from left to right across the periodic table.** This trend occurs because the increasing nuclear charge generates an increased hold on the valence electrons.
- **Within groups, ionization energy tends to *decrease* from top to bottom down the periodic table.** This trend occurs because atomic size increases as you move down a group. As a result, less energy is needed to remove an electron that is farther from the nucleus.

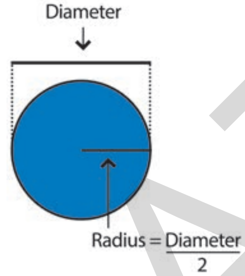


*In the table above, ionization energy is measured in kilojoules per mol (kJ/mol).
**The ionization energy is unknown for the elements on the table shown in white.
***The arrows indicate the *general trend* of ionization energy. It increases as you move to the right within a period and as you move from the bottom to the top within a group.

As you move to the right on the periodic table, the atoms are getting closer to having a full octet, which will make them more stable. Since atoms want to be stable, elements like the halogens (fluorine, chlorine, etc.) that are only one electron away from having a full octet will want to *gain* an electron, so it will take more energy to take one away. On the other hand, elements like sodium that need to *lose* an electron to reach a full octet will have a lower ionization energy. Notice that noble gases have higher ionization energies because they already have a full octet and are therefore stable.

ATOMIC RADII

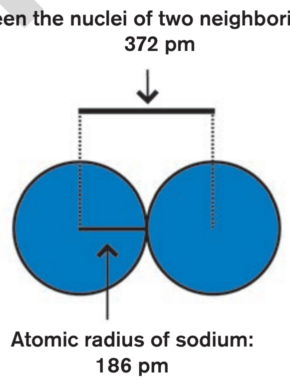
The **atomic radius** of an element's atom is a measure of atomic size. For a single atom, it can be considered as the typical distance from the nucleus to the boundary of the electron cloud. Think of an atom as a ball. You can find the diameter of the ball by measuring from one edge of the ball to the other edge. The radius can be found by measuring from the center of the ball to the edge. It can also be found by dividing the diameter by two.



An atom does not have a clearly defined edge because of the electron cloud. The atomic radius is determined by how close one atom is to a neighboring atom. It is half the distance between the nuclei of two atoms of an element. This distance is so small that it is typically measured in picometers (10^{-12} m).

EXAMPLE:

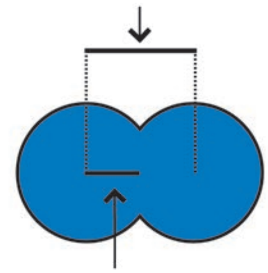
The atomic radius for bonded metallic atoms in an elemental sample, like sodium, is half the distance between the nuclei of the two neighboring atoms.



EXAMPLE:

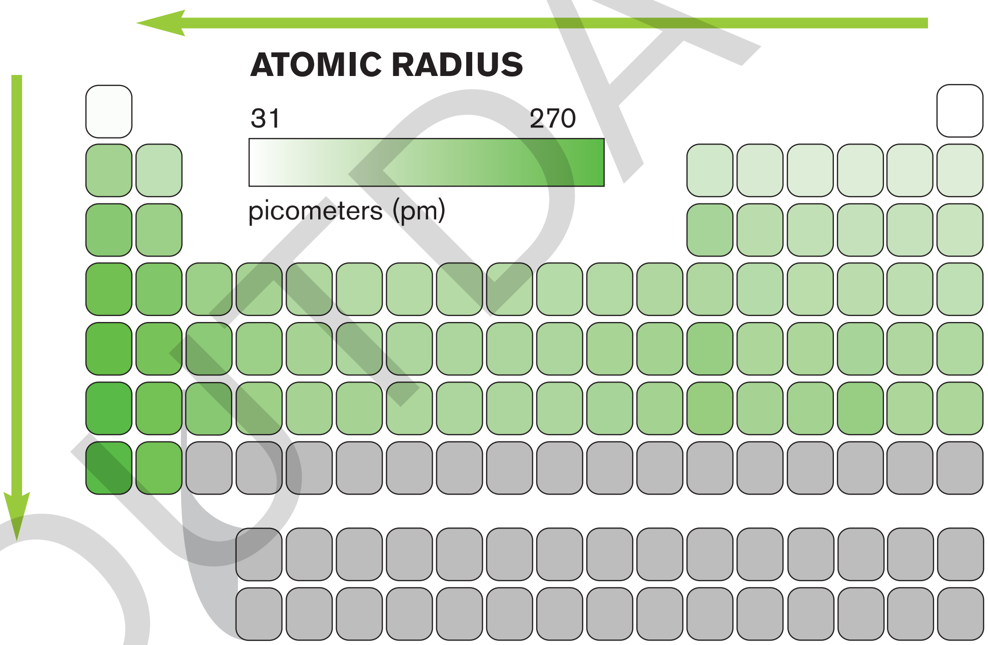
In a bonded nonmetal, like oxygen and other diatomic molecules, the radius is half the distance between the nuclei of the atoms in the molecule.

Distance between the nuclei of oxygen atoms joined in a diatomic oxygen molecule:
146 pm



Atomic radius of oxygen:
73 pm

- **The size of atomic radii tends to *decrease* from left to right across the periodic table.** This trend results from the increasing positive charge in the nucleus. The valence electrons are also less shielded from the nucleus's charge, so the nucleus can pull the outer electrons inward.
- **The size of atomic radii tends to *increase* as you move down a group of the periodic table.** This trend is caused by the increasing energy level (and thus larger orbitals). The electrons are farther from the pull of the positive nucleus.



*In the table above, the atomic radii are measured in picometers (pm).
**The atomic radius is unknown for the elements on the table shown in gray.
***The arrows indicate the *general trend* for atomic radii. It increases as you move to the *left* within a period and as you move from the *top to the bottom* within a group.

Knowing about the atomic radius of an element can help you think about its ionization energy. If the radius of an element is small, this tells you that the electrons are closer to the nucleus. This means that the protons in the nucleus are pulling in the electrons on the outer shells – there is a stronger attraction. If you want to pull an electron off (which takes ionization energy), it will take a lot more energy for a small atom as opposed to an atom that has many, many more electrons between the valence electrons and the nucleus. When there are more electrons, the attraction between the nucleus and the outermost electrons is weaker, so pulling off an electron will be easier.

CHEMICAL BONDS

A **chemical bond** is an intramolecular force of attraction that holds together atoms in a molecule or compound. Bonds are formed as a result of the attraction between the positive nucleus of one atom and the negative electrons of another atom. Chemical bonds can also be formed by the attraction between positive and negative ions.

Atoms work to achieve a stable octet arrangement of valence electrons. Atoms and molecules give, take, or share their valence electrons during chemical reactions in order to reach this stable state. When atoms or ions of different elements interact, chemical bonds can be formed, broken, or rearranged to create new compounds. Therefore, 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 (H_2). Likewise, a molecule of oxygen gas (O_2) 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 the Chemical Reactions section).

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

TYPES OF CHEMICAL BONDS

There are three types of chemical bonding: ionic, covalent, and metallic. Bonding involves only the valence electrons of an atom.

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, so 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 positive ion (a cation). Since the atoms have opposite charges and become attracted to one another, this force of attraction holds the atoms together. Compounds held together by ionic bonds are called ionic compounds.

Ionic bonds typically occur between atoms of elements located on opposite sides of the periodic table.

EXAMPLE:

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

Step 1: Na → Na^+ + electron (production of an Na cation plus release of electron)

Step 2: electron + Cl → Cl^- (released Na electron reacts with Cl to produce a Cl anion)

Combined: Na + ~~electron~~ + Cl → Na^+ + Cl^- + ~~electron~~

Notice that the electron produced in Step 1 is used in Step 2, so it is cancelled out in the combined reaction. (See the section on Ions from *The Core of Chemistry*.)

EXAMPLE:

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

Step 1: Mg → Mg^{2+} + 2 electrons (production of an Mg cation plus release of electrons)

Step 2: 2 electrons + 2 Cl → 2 Cl^- (released Mg electrons react with Cl to produce Cl anions)

Combined: Mg + ~~2 electrons~~ + 2 Cl → Mg^{2+} + 2 Cl^- + ~~2 electrons~~

Notice that twice as many Cl atoms are needed to take in the two electrons released by the Mg, making the chemical formula $MgCl_2$.

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. **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

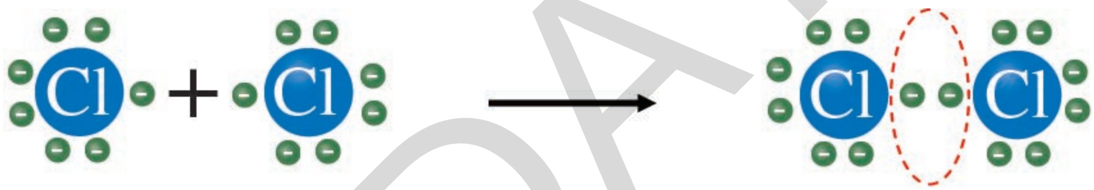
Covalent bonds occur when valence electrons are *shared* between two nearby atoms. Compounds formed from atoms that share electrons through a covalent bond are called covalent compounds.

- Covalent bonds create stable compounds if the sharing of electrons brings about a noble gas configuration for each atom (with eight 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:

Look at the gaseous molecule Cl_2. Chlorine is found on the third period of the periodic table and is part of the halogen group. Because this molecule is made of two chlorine atoms, each of the atoms equally wants to be configured like the nearest noble gas element, argon. The two chlorine atoms agree to share each other's electrons.

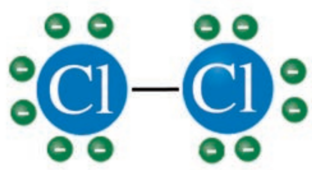
Remember that only the outermost electrons participate in the bond. Therefore, the seven valence electrons in the third energy level participate in the bond. The reaction can be shown as:



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

A single covalent bond contains two electrons. This means there is one covalent bond in a Cl_2 molecule (2 electrons shared, divided by 2 electrons in each bond = 1 bond).

- Two electrons that form a bond are shown as a line. Cl_2 is shown as:



- The bond that forms between the chlorine molecules to make Cl_2 is called a *single covalent bond*. Other chemicals may contain *double 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. Covalent bonds typcally occur between nonmetal elements.

Quick Fact

Remember, these structures that show atomic centers (symbols) and either lines or dots for the valence electrons are called Lewis structures.

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

Think About It...

Why must atoms have similar electronegativities to form covalent bonds? What happens when atoms have very different electronegativities?

METALLIC BONDING

Metallic bonding occurs when the atoms in a metal contribute their electrons to a "sea" of shared electrons. This "sea" of electrons spans the entire structure.

- Metallic bonds are collective by nature, so a single metallic bond does not exist.
- In a metal, the valence 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 metal atom gives up its valence electrons, forming a "sea" of electrons.
- Elements along the left side of the periodic table often form metallic bonds.
- Metallic bonds also form among elements that have 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. For example, many atoms of gold will come together to form strong metallic bonds that will not give up electrons to other substances.
- Some metallic elements are actually quite easy to keep in pure form because they are relatively unreactive. These elements include gold, copper, and silver.

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.

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.

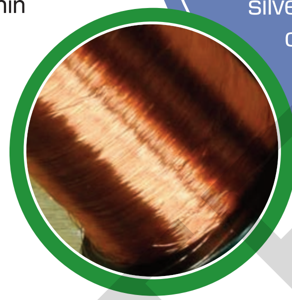
Think About It...

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?

- **Malleability:** the ability of a metal to be flattened, shaped, or formed, without breaking, when pressure is applied. This includes the ability of a metal to be hammered into a thin sheet.
  - The mobility of electrons allows metal atoms to slide past one another when stress is applied. They do so without experiencing strong repulsive forces that would cause other materials to shatter.
- **Ductility:** the ability of a metal to be stretched into a thin wire or thread without breaking.
  - Like malleability, the mobility of electrons in a metallic bond allows the atoms to slide past one another as the metal is pulled and reshaped.

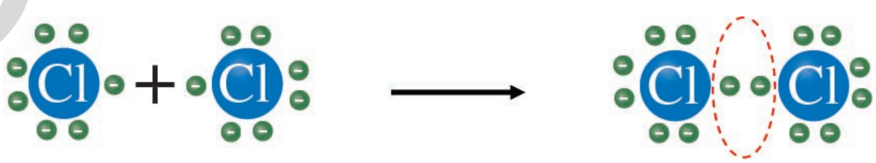
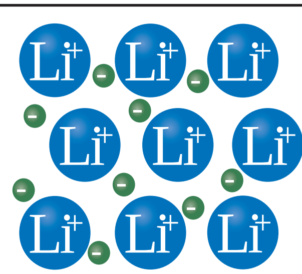
Quick Fact

Gold, silver, and copper (all group 11 elements) are highly malleable and ductile. They are also good conductors of electricity. Because of these properties, gold, silver, or copper could be used to make electrical wires. However, gold and silver are expensive metals, so copper is commonly used in wiring.



BONDING REVIEW

- Ionic bonding is essentially the result of an atom donating an electron to another atom so that they both complete their octets.
- Covalent bonding is the result of atoms that both need electrons, so they share.
- Metallic bonding is the result of collectively shared electrons.

Ionic Bonding	
Covalent Bonding	
Metallic Bonding	

GRAVITY

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. On the earth, all objects are pulled toward the earth's center. If you slide a book over the edge of a table or let go of a bag you are holding, you can expect both objects to drop to 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.

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 the Measurement in *The Core of Chemistry*).

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. For example, as a spacecraft leaves the earth's atmosphere, its attraction to the earth decreases as it moves farther away.

HISTORY: **ISAAC NEWTON** (1643-1727)

Sir Isaac Newton was 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 proposed 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 today's large telescopes.

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 paper clips, nails, or hairpins, will stick to a type of object called a magnet? A **magnet** is an object that creates a strong magnetic field (an area of magnetic force). Only certain metals produce a magnetic field, including iron, nickel, and cobalt. All magnets are made of these types of 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. Two like magnetic poles will repel each other. For example, 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). Thus, metals with magnetic properties will be attracted to a magnet. Specifically, the north pole 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.

Quick Fact

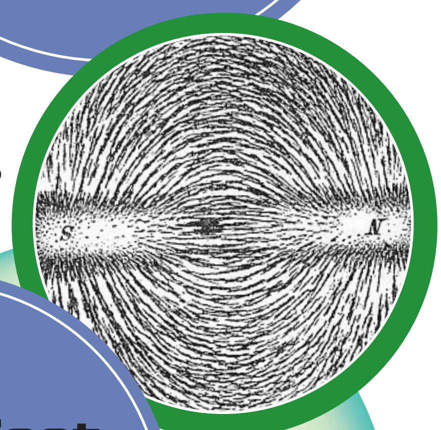
If you hang a magnet by a string, its north pole will turn North. Therefore, the north pole is sometimes called the north-seeking pole. Likewise, the south pole will point South.

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.

Quick Fact

Iron filings are small pieces of iron that together look like a powder. 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.

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, an object like an automobile that is lifted or moved by a magnet at a junkyard acts as a temporary magnet. That object generally loses it magnetism when the permanent magnet is removed. However, in some cases, the magnetized object will still have weak magnetic properties.

Quick Fact

Another main type of magnet is an electromagnet. The magnetic field in these magnets is produced by an electric current.

HISTORY: **MAGNETIC MATERIALS**

Humans have known about magnetic materials for over 2,000 years. The ancient Greeks discovered a unique type of rock in the city of Magnesia. These rocks attracted materials containing iron. They 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. That way, one end of the pointer will always point toward the North.

NOTES

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. Sometimes one or more of the products can be classified as byproducts. A **byproduct** is a product that is created at the same time as the primary product(s).

In general, a chemical reaction will be represented like this:

$$\text{Reactants} \rightarrow \text{Products}$$

EXAMPLE:
The simple chemical reaction between hydrogen and nitrogen is shown below:

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

- **The hydrogen (H_2) and nitrogen (N_2) molecules are the reactants; the resulting ammonia (NH_3) is the product.**
- **Heat energy initiates the reaction.**

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).

$$A + B \rightarrow AB$$

EXAMPLE:

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

The (g) and (l) after the compounds correspond to the phases that those compounds are in. A (g) means that compound is a gas, (l) means a liquid, and an (s) tells you it is a solid. You might also see (aq), which stands for aqueous solution – or a mixture of a compound where water is the solvent.

Think About It...

The reaction to the left could ideally be used for hydrogen-powered cars. For this reaction to power a hydrogen car, there is plenty of oxygen in the air, but how would 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.

$$AB \rightarrow A + B$$

EXAMPLE:

$$2HgO\ (s) \xrightarrow{\Delta} 2Hg\ (l) + O_2\ (g)$$

In this decomposition reaction, mercury oxide (HgO) splits into mercury and oxygen gas. The small triangle above the arrow means the reaction needs energy to take place.

Quick Fact

If you wanted to say what was occurring in the reaction (left) you would describe it by saying: "Two mercury oxide molecules decompose into two mercury atoms plus one oxygen gas molecule."

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.

$$A + BC \rightarrow AC + B$$

EXAMPLE:

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

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).

Think About It...

How would you describe the displacement reaction to the left?

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

$$AB + CD \rightarrow AD + CB$$

EXAMPLE:

$$Fe_2O_3\ (s) + 6HCl\ (aq) \rightarrow 2FeCl_3\ (aq) + 3H_2O\ (l)$$

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

Quick Fact

The reaction to the left could also be written as:
$AB + CD \rightarrow DA + BC$

A molecule composed of a single D atom and a single A atom can be written as either DA or AD, but scientists have adopted a rule that says to place the more electropositive (least electronegative) element first. 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 the result of an atom or group of atoms gaining or losing electrons.

- **Oxidation**: a chemical reaction that involves an increase in oxidation number. Oxidation results from an atom or group of atoms losing one or more electrons.
- **Reduction:** a chemical reaction that involves a decrease in oxidation number. Reduction 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. Some elements are *oxidized* while others are *reduced*, but the number of electrons on both sides of the equation remains the same.

EXAMPLE:

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

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*: The sodium atoms begin with an oxidation number of zero and end with an oxidation number of +1. They have been oxidized from sodium atoms to sodium cations.**

$Na \rightarrow Na^+$ + electron (production of an Na cation plus release of an electron)

***Reduction*: The chlorine gas (Cl_2) begins with an oxidation number of zero and ends with an oxidation number of -1. The chlorine atoms have been reduced to chloride anions.**

electron + $Cl \rightarrow Cl^-$ (released Na electron reacts with Cl to produce a Cl anion)

In oxidation-reduction reactions, the atom that is reduced is called the oxidizing agent, and the atom that is oxidized is called the reducing agent. In the reaction above, chlorine oxidizes the sodium atoms so it is the *oxidizing agent*. On the other hand, sodium reduces the chlorine atoms, so 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.

Some of 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 until a stable product is formed. 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 product 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, one or more of the starting materials are exhausted (used up).

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

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.

EXAMPLE:

Initiation:	Cl_2 + light → Cl· + Cl·
Propagation:	Cl· + H_2 → HCl + H·
	H· + Cl_2 → HCl + Cl·
	Cl· + H_2 → HCl + H·
	H· + Cl_2 → HCl + Cl·
Termination:	Cl· + Cl· → Cl_2

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 separated chlorine atoms, called chlorine free radicals (Cl•).**
- **The chlorine radicals are very reactive because they do not have eight electrons in their outer valence shell. Therefore, they 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. In this case, this reaction occurs when two chlorine radicals combine with each other, to form Cl_2, which is a stable molecule known as chlorine gas.**

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:

$$A + B \rightleftharpoons C + D$$

EXAMPLE:

$$2NaCl + CaCO_3 \rightleftharpoons Na_2CO_3 + CaCl_2$$

The forward reaction can be seen in saltwater lakes:

$$2NaCl + CaCO_3 \rightarrow Na_2CO_3 + CaCl_2$$

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 is as follows:

$$Na_2CO_3 + CaCl_2 \rightarrow 2NaCl + CaCO_3$$

Sodium carbonate reacts with calcium chloride to produce sodium chloride and calcium carbonate. Both reactions coexist in nature and are happening at the same time. If only the forward reaction occurred, limestone rocks would quickly dissolve in saltwater lakes – something that does not easily occur in nature.

- 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 concentrations of the reactants and products does not change when conditions remain the same. However, the concentrations of the reactants and products do not have to be equal to each other.

Equilibrium describes how far a reaction goes. For instance, it describes how much product a reaction can produce (unless we manipulate it!).

Quick Fact

Scientists often try to manipulate equilibrium.

Doing this can cause a particular reaction to make more product than usual.

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 produce/release energy.
  - Exothermic reactions may occur spontaneously. They often release energy in the form of heat, light, or sound.

$$A + B \rightarrow C + D + \text{energy}$$

EXAMPLE:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + \text{energy}$$

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

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

$$A + B + \text{energy} \rightarrow C + D$$

EXAMPLE:

$$\text{energy} + 2Al_2O_3 \rightarrow 4Al + 3O_2$$

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

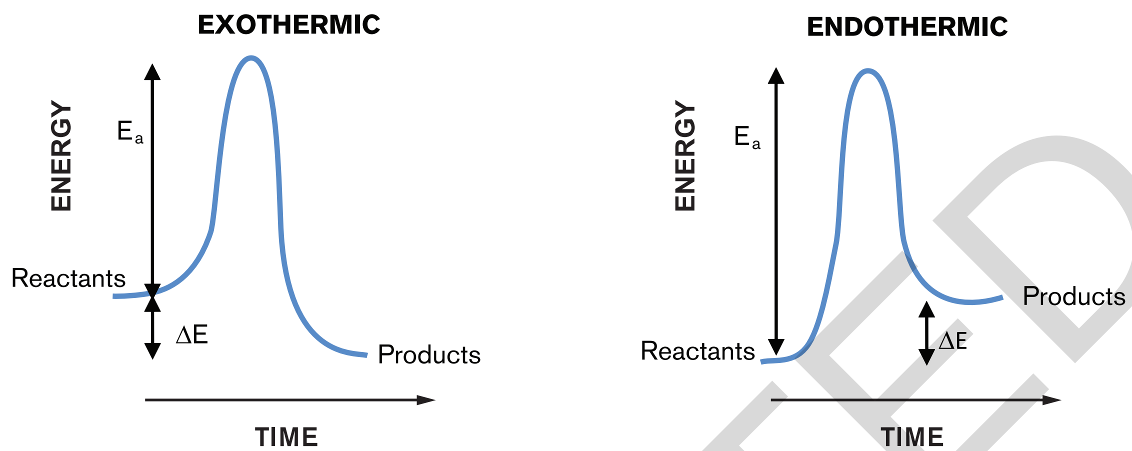
- **Energy of activation:** the amount of energy needed to cause a chemical reaction.
  - Energy of activation (activation energy) is represented by E_a.
  - As the diagrams on the following page illustrate, the activation energy is the energy required to make the reaction occur. It is the energy needed 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 the products 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 usually increases the rate of the reaction.
- In a reversible reaction that is at equilibrium, an increase in the concentration of the products typically decreases the rate of the reaction.

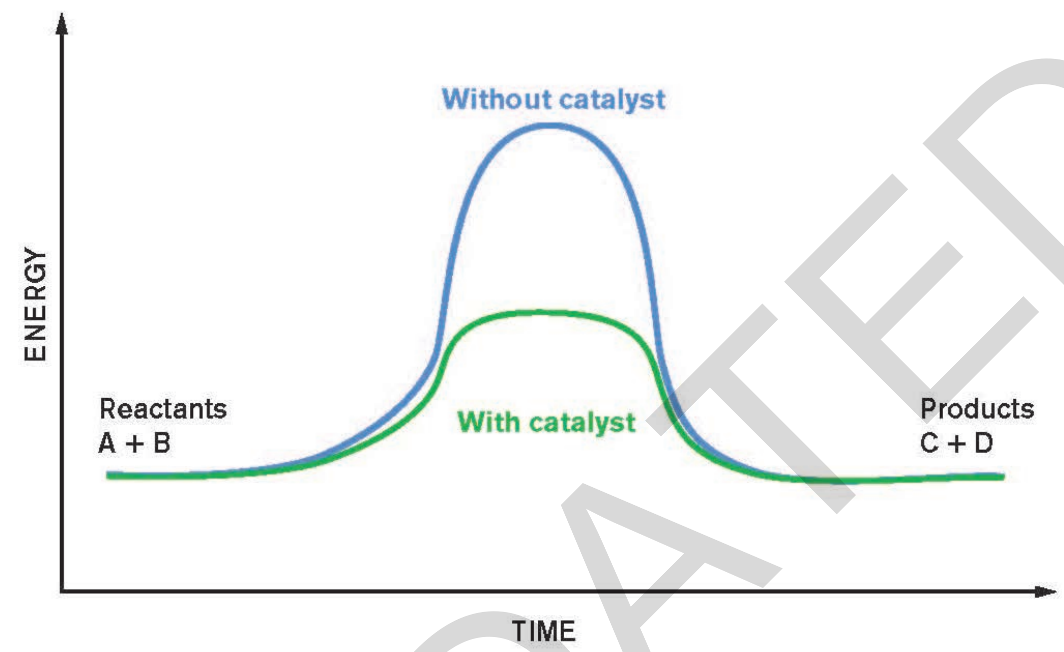
A **catalyst** is a substance that allows a chemical reaction to take place at a different rate or under different conditions. During the reaction, the catalyst is not consumed or changed.

- The catalyst usually changes the pathway the reaction takes.
- The new pathway usually involves a lower energy barrier. This lower barrier often 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. Your body doesn't require the amount of heat that a car needs to burn fuel, but it does require a lot of energy. However, you have special catalysts in your body called enzymes, which allow you to burn fuel at normal body temperature.

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



Because of the lower energy barrier (or lower activation energy), catalysts can help make a reaction occur faster. However, that is not always the case.

Although catalysts can change the speed of the reaction, they don't change the energy of the products or reactants. From the graph above, notice that the energies of the reactants and the products remain unchanged.

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**, which uses chemical symbols and formulas to describe the reaction. A chemical equation shows the reaction that occurs using the chemical symbols of the compounds involved.

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 changed. 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. Therefore, 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:

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

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:

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

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 $1C_1H_4$.
- To make sure the equation 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:

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

Reactant Side of Equation	Product Side of Equation
C: $1 \times 1 = 1$	C: $1 \times 1 = 1$
H: $1 \times 4 = 4$	H: $2 \times 2 = 4$
O: $2 \times 2 = 4$	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. His law stated 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.

BALANCING CHEMICAL EQUATIONS

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

$$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$$

- 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 so 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 1.

$$C_2H_6 + O_2 \rightarrow 2CO_2 + H_2O$$

- 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***. Remember that the subscripts tell you how many atoms of each are needed in the molecule. In this case, you need 1 carbon atoms and 2 oxygen atoms to make 1 molecule of CO_2. The coefficient tells you that you make 2 molecules of CO_2.
- Add a coefficient of 3 in front of H_2O on the right side, making it $3H_2O$.

$$C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$$

- 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 (2 on the left side and 7 on the right side).
- Add a coefficient of 3.5 in front of O_2 on the left side, making it $3.5O_2$.

$$C_2H_6 + 3.5O_2 \rightarrow 2CO_2 + 3H_2O$$

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

Quick Fact

To balance a chemical equation, it is usually easiest to balance the elements that appear in the fewest chemical formulas. For example, in the equation shown, it is easiest to begin with carbon and hydrogen. They appear twice, while oxygen appears three times.

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.

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

This equation is the properly balanced equation for the reaction.

SECTION V:
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 of a solution measures how acidic or basic it is. 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. When a solution has a lot of hydrogen ions that are released into the water (high concentration of hydrogen ions), it is very acidic.
- 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. A substance with a pH of 1 is a very strong acid.
  - The more basic a substance is, the higher its number on the pH scale. 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. Water has an equal number of hydrogen (H^+) ions and hydroxide (OH^-) ions, giving it a neutral pH.

$$H^+ + OH^- \rightarrow H_2O$$

  - Negative pH: some very strong acids may have a pH lower 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	
Stomach acid	1.0–2.0	
Lemon juice, cola	2.3–2.5	
Vinegar	2.9	
Apple juice, orange juice	3.3–3.8	
Coffee	5.0–5.5	
Milk	6.5	
Pure water	7.0	
Human blood	7.4	
Sea water	8.0	
Baking soda solution	8.5–9.0	
Milk of magnesia	10.5	
Household ammonia	11.5–12.0	
Bleach	12.5	
Liquid drain cleaner	13.5–14.0	

*See the subsection on Indicators.

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

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

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*.

EXAMPLE:

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

Ammonia is a base but *not* an alkali.

Quick Fact

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

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 reveal slightly different colors depending on the strength of the acid (see the subsection on Indicators).

The chemical equation of an acid dissociating (producing hydrogen ions) looks like this:

$$HA\ (aq) \rightarrow A^-\ (aq) + H^+\ (aq)$$

Many things will affect the strength of an acid or a base:

- For some acids, the strength of an acid may be affected by the size of the anion (see the 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. Hydroiodic acid ($HI_{(aq)}$) is a stronger acid then hydrofluoric acid ($HF_{(aq)}$).
  - Electronegativity also affects how strong an acid is. A more electronegative atom (like fluorine compared to iodine) will be tightly bonded with a hydrogen cation (H^+) so that its octet can be complete, meaning there will not be as many H^+ ions dissociated in the solution. (See the subsection on Electronegativity.)
- 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.

$$BOH(aq) \rightarrow B^+(aq) + OH^-(aq)$$

- 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

SECTION VI:
RADIOACTIVITY & NUCLEAR REACTIONS

OBJECTIVES

- Define radioactivity and radioisotopes.
- Explain half-life and use it in calculations.
- Identify common radioactive elements and describe their properties.
- Describe the difference between nuclear fission and nuclear fusion.
- Identify man-made elements and their location on the periodic table.

Elements tend to exist in more than one form, called isotopes. Isotopes of elements differ in the number of neutrons in their nucleus, so they also differ in their mass numbers (see the section on Isotopes from *The Core of Chemistry*).

Isotopes of an element have similar chemical properties to each other and undergo similar reactions. However, since isotopes differ in atomic mass, their physical properties are not exactly the same. Isotopes of an element often undergo chemical reactions at different rates.

EXAMPLE:

Hydrogen is composed of three isotopes (protium, deuterium, and tritium). The lightest isotope, protium, tends to undergo chemical reactions at a faster rate.

Quick Fact

Carbon-12 is the most common form of carbon. It was adopted in 1961 as the standard for defining all atomic weights.

Carbon-13 is non-radioactive and is frequently used for isotopic labeling studies. These studies follow how a carbon atom goes through specific reactions.

Carbon-14 is used in a process called carbon dating.

– After 5,730 years, half of the nuclei in a sample of a carbon-14 decay. (This period of time is its half-life.)

– Scientists use the predictable decay of carbon-14 to determine the age of organic materials up to 50,000 years old.

– Carbon dating is useful for studying artifacts left behind by ancient cultures.

RADIOACTIVITY

Radioactivity is the spontaneous breakdown of an unstable nucleus in an atom that involves the release of energy in the form of electromagnetic radiation or particles.

- **Radioisotopes**: atoms that are radioactive.

The **half-life** of an isotope is the time it takes for one-half of the nuclei present in a sample to undergo radioactive decay.

- After one half-life, 50% of the original sample will remain.
- After two half-lives, 25% of the original sample will remain, and so on.

HISTORY: **HENRI BECQUEREL** (1852–1908)

In 1896, French physicist Henri Becquerel accidentally discovered radioactivity. He was actually investigating phosphorescence in uranium salts. In 1903, he shared the Nobel Prize in physics with Pierre and Marie Curie in recognition of his discovery and their study of natural radioactivity.

The SI unit for radioactivity, the becquerel (Bq), is named after him.

HISTORY: **MARIE CURIE** (1867–1934)

Marie Curie discovered that the element thorium was "radioactive," a term she created. The same year, a German scientist named Gerhard Schimdt also made the same discovery about thorium.

Curie, along with her husband Pierre, discovered the radioactive elements polonium and radium.

In 1903, the Curies and Henri Becquerel were awarded the Nobel Prize in physics for the discovery and exploration of natural radioactivity. In 1911, Curie received her second Nobel Prize in chemistry for isolating radium and determining its atomic weight. She was the first woman to receive a Nobel Prize and the only woman, to this day, to receive two Nobel Prizes.

Po

POLONIUM

Atomic #84

Polonium was discovered by Marie and Pierre Curie in 1898. Therefore, the element was named after the country Poland, where Marie Curie was born.

Characteristics:

- Is a very rare natural element, found in extremely small amounts in uranium ores.
- Is mainly used as a source of neutrons, generally by combining it with beryllium.
- Has specialty uses in eliminating static electricity in machinery and removing dust from photographic film.

Polonium has over 25 known isotopes. Its most common isotope, Po-210, has a half-life of only 138 days. The radioactive decay of Po-210 produces a lot of heat (140 watts per gram).

Rn

RADON

Atomic #86

Radon comes from the radioactive decay of the element radium.

Characteristics:

- Is radioactive; the isotope with the longest half-life is radon-222 with a half-life of only four days.
- Is a colorless radioactive gas at a normal room temperature of about 70–75 °F.
- Glows with a yellow color when cooled to its solid state.
- Is emitted naturally, in some regions, from the soil and rocks and can sometimes build up in people's homes.

The World Health Organization estimates that 15% of all lung cancer cases are caused by exposure to radon. Radon test kits are available to check for radon accumulation in homes, especially basement levels.

Ra

RADIUM

Atomic #88

Radium was discovered by Marie and Pierre Curie in 1898. Its name comes from the Latin word "radius" meaning "ray."

Characteristics:

- Is a highly reactive metal.
- Is a brilliant, white metal in pure form but blackens when exposed to air.
- Occurs naturally in the environment from the decay of uranium and thorium.

Its most stable isotope, radium-226, has a half-life of about 1,600 years.

Pure radium and some of its compounds glow in the dark. As a result, radium was used in the mid 1900s in a luminous paint on the hands and numbers of watches to make them glow in the dark. However, this practice stopped when the risks of radium exposure became known.

The radioactive decay of an unstable nucleus may release several types of radiation, including alpha radiation, beta radiation, and gamma radiation.

Quick Fact
Radiation may be used in medicine to treat disease and to look inside the body to diagnose medical problems. Radiation has proven useful to kill cancer cells by causing mutations (defects) in DNA, thus preventing the cancerous cells from being able to grow and divide.

ALPHA (α) RADIATION (ALPHA PARTICLES): radiation composed of helium-4 nuclei (having a nucleus that is the same as helium with two protons and two neutrons).

- Alpha radiation travels very slowly and only a very short distance through air. It cannot penetrate skin or even a thin sheet of paper.
- Alpha particles are not radioactive. After losing their energy, they attract two electrons to become a helium atom.

BETA (β) RADIATION (BETA PARTICLES): radiation composed of electrons, emitted from an unstable nucleus, that are in high velocity.

- Beta radiation can travel several meters through air but is stopped by solid materials.
- Beta particles can penetrate human skin, but clothing often helps to block most beta particles.
- If the release of a beta particle does not get rid of the extra energy in an unstable nucleus, the nucleus will often release the rest of the excess energy in the form of gamma rays.

GAMMA (γ) RADIATION (GAMMA RAYS): radiation composed of high-energy photons in the form of electromagnetic radiation.

- Gamma radiation is able to travel many meters in air. It easily penetrates most materials, including several centimeters through human tissue.
- Gamma radiation frequently accompanies the emission of alpha and beta radiation.

HISTORY: JOHANNES WILHELM GEIGER (1882–1945)

Johannes Wilhelm "Hans" Geiger was a German physicist known for his work on radioactivity. In 1928, with fellow physicist Walther Müller, he developed a device to measure radioactive emissions. The device became known as the **Geiger Counter**. The two worked to improve the device's sensitivity, performance, and durability. As a result, they created a tool that is used in laboratories around the world today.

NUCLEAR ENERGY

Nuclear reactions are changes that occur in the structure of atomic nuclei. The energy that results from nuclear reactions is called nuclear energy or atomic energy. Nuclear energy is released from atoms in two different ways: nuclear fission and nuclear fusion.

NUCLEAR FISSION: a nuclear reaction that occurs when an atomic nucleus splits into two smaller parts (nuclei), usually about the same size. When this happens, energy is released.

- Uranium nuclei can be split easily by bombarding them with neutrons.
- Once a uranium nucleus is split, multiple neutrons are released, initiating a chain reaction as other uranium nuclei are split.

U

URANIUM

Atomic #92

Uranium was first identified in pitchblende ore in 1789. It was named after the planet Uranus, which had been discovered around that time.

Characteristics:

- Is the heaviest naturally occurring element on Earth, except for minute traces of neptunium and plutonium.
- Is highly radioactive, toxic, and carcinogenic.
- Has over 16 isotopes, all of which are radioactive.

Uranium's radioactivity was first detected by Henri Becquerel in 1896. Today, it is primarily used in nuclear fuels and explosives. Uranium, specifically the isotope uranium-235, is the principle element used in nuclear reactors and in certain types of atomic bombs.

Uranium compounds have been used for centuries as additives in glass. They give the glass interesting yellow and green colors and fluorescent effects.

NUCLEAR FUSION: a nuclear reaction that occurs when the nuclei of atoms join to make a larger nucleus. Again, energy is given off in this reaction.

- Nuclear fusion only occurs under very hot conditions.
- The sun and all other stars create heat and light through nuclear fusion. In the sun, hydrogen nuclei fuse to make helium.

Quick Fact

The hydrogen bomb uses nuclear fusion. Hydrogen nuclei fuse to form helium. In the process, they release huge amounts of energy and create a massive explosion.

HISTORY: **ENRICO FERMI** (1901–1954)

Enrico Fermi was an Italian physicist most noted for his work on beta decay, for the development of the first nuclear reactor, and for his contributions to the development of quantum theory. He worked on the Manhattan Project during World War II to produce the atomic bomb, though he warned of its power.

In 1938, Fermi won the Nobel Prize in physics for his work on radioactivity. Element 100 was named after him.

Pu

PLUTONIUM

Atomic #94

Plutonium was discovered in 1940 at the University of California at Berkeley. It was named after the dwarf planet Pluto. The element's discovery, however, was kept classified by the government until 1946.

Characteristics:

- Is a very heavy, silvery metal in pure form.
- Is a rare radioactive element; found in minute amounts (one part per trillion) in uranium ore.
- Used mainly as a fuel for nuclear reactors and nuclear bombs.

Plutonium was produced in large quantities in the U.S. during World War II, as part of the Manhattan Project to create the atomic bomb.

Over one-third of the energy produced in most nuclear power plants comes from plutonium.

NOTES

MAN-MADE ELEMENTS

The elements above atomic number 92 are known as transuranic or transuranium elements. They do not occur naturally on the earth. Most of these heavier elements have been made by bombarding the element uranium with neutrons or other particles in a cyclotron.

Many of the man-made, transuranic elements are named for important chemists or physicists:

Atomic #	Element	Symbol	Named for ...
99	Einsteinium	Es	Albert Einstein, the famous scientist who developed the Theory of Relativity.
101	Mendelevium	Md	Dmitri Mendeleev, who developed the periodic table.
102	Nobelium	No	Alfred Nobel, who commercialized dynamite and endowed the Nobel Prizes for physics, chemistry, medicine, literature, and peace.
103	Lawrencium	Lr	Ernest O. Lawrence, who invented the cyclotron.
104	Rutherfordium	Rf	Ernest Rutherford, who helped develop the modern understanding of the atomic nucleus.
106	Seaborgium	Sg	Glenn Seaborg, who was known for his work in the separation and purification of plutonium. He was also known for proposing the "Actinide" concept for reorganizing the periodic table.
107	Bohrium	Bh	Niels Bohr, who proposed a model of atomic structure that explained the role of the electron.
109	Meitnerium	Mt	Lise Meitner, who co-discovered nuclear fission.

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