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CHEMISTRY CONNECTIONS



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SECTION I: CHEMICAL FORMULAS

OBJECTIVES

- Identify the Lewis symbols 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.

LEWIS SYMBOLS

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 symbols**. Lewis symbols (also known as electron dot structures) contain the element's chemical symbol and dots that represent the high-energy outermost electrons, called valence electrons. Remember, each element has a unique chemical symbol of one or two letters, as shown on the periodic table.

VALENCE ELECTRONS

Valence electrons are the electrons in the highest energy level, located in the electron shell the farthest from the nucleus of an atom (see **Periodic Trends**). Atoms often react using their valence electrons, so looking at a Lewis symbol and knowing how many valence electrons an element has can help determine how it will react with other elements. The electrons not in the highest energy level are known as **core electrons**, and are not usually involved in chemical reactions.

Most main group elements (elements in groups 1-2 and groups 13-18) can have up to eight valence electrons. Transition metals (groups 3-12) do not follow this rule.

Groups 1-2 and 13-18 (the main group elements) all follow the same pattern of valence electrons. Moving from left to right across the periodic table, one electron is added with each group. The number of valence electrons increases by one with each group from left to right, not including the transition metals (groups 3-12, colored purple), as shown on the following page.



Quick Fact

The ground-state electron configuration of an atom shows the lowest energy state of the atom. In the presence of light, electrons can sometimes absorb energy to jump to a higher energy level, changing the electron configuration. This is called an **excited state**.





| Valence electrons | 1 | 2 | | | | | | | | | | | 3 | 4 | 5 | 6 | 7 | 8 |
|-------------------|----------------------------------|----------------------------------|--------------------------------------|-------------------------------------|----------------------------------|----------------------------------|----------------------------------|-----------------------------------|----------------------------------|------------------------------------|-----------------------------------|-----------------------------------|-----------------------------------|----------------------------------|-----------------------------------|-----------------------------------|----------------------------------|---------------------------------|
| Group | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
| 1 | H Hydrogen 1.00794 | | | | | | | | | | | | | | | | | He Helium 4.0026 |
| 2 | Li Lithium 6.941 | Be Beryllium 9.0122 | | | | | | | | | | | B Boron 10.811 | C Carbon 12.011 | N Nitrogen 14.0067 | O Oxygen 15.9994 | F Fluorine 18.9984 | Ne Neon 20.183 |
| 3 | Na Sodium 22.9898 | Mg Magnesium 24.305 | | | | | | | | | | | Al Aluminum 26.9815 | Si Silicon 28.086 | P Phosphorus 30.9738 | S Sulfur 32.066 | Cl Chlorine 35.453 | Ar Argon 39.948 |
| 4 | K Potassium 39.098 | Ca Calcium 40.08 | Sc Scandium 44.956 | Ti Titanium 47.87 | V Vanadium 50.942 | Cr Chromium 51.996 | Mn Manganese 54.938 | Fe Iron 55.845 | Co Cobalt 58.9332 | Ni Nickel 58.69 | Cu Copper 63.546 | Zn Zinc 65.39 | Ga Gallium 69.72 | Ge Germanium 72.61 | As Arsenic 74.9216 | Se Selenium 78.96 | Br Bromine 79.904 | Kr Krypton 83.80 |
| 5 | Rb Rubidium 85.47 | Sr Strontium 87.62 | Y Yttrium 88.906 | Zr Zirconium 91.22 | Nb Niobium 92.906 | Mo Molybdenum 95.94 | Tc Technetium (98) | Ru Ruthenium 101.07 | Rh Rhodium 102.905 | Pd Palladium 106.4 | Ag Silver 107.868 | Cd Cadmium 112.41 | In Indium 114.82 | Sn Tin 118.71 | Sb Antimony 121.76 | Te Tellurium 127.60 | I Iodine 126.9045 | Xe Xenon 131.29 |
| 6 | Cs Cesium 132.905 | Ba Barium 137.33 | 57-71 Lanthanides | Hf Hafnium 178.49 | Ta Tantalum 180.948 | W Tungsten 183.84 | Re Rhenium 186.2 | Os Osmium 190.2 | Ir Iridium 192.2 | Pt Platinum 195.08 | Au Gold 196.967 | Hg Mercury 200.59 | Tl Thallium 204.38 | Pb Lead 207.2 | Bi Bismuth 208.98 | Po Polonium (210) | At Astatine (210) | Rn Radon (222) |
| 7 | Fr Francium (223) | Ra Radium (226) | 89-103 Actinides | Rf Rutherfordium (261) | Db Dubnium (268) | Sg Seaborgium (271) | Bh Bohrium (272) | Hs Hassium (277) | Mt Meitnerium (276) | Ds Darmstadtium (281) | Rg Roentgenium (280) | Cn Copernicium (285) | Nh Nihonium (284) | Fl Flerovium (289) | Mc Moscovium (288) | Lv Livermorium (293) | Ts Tennessine (294) | Og Oganesson (294) |
| | La Lanthanum 138.91 | Ce Cerium 140.12 | Pr Praseodymium 140.908 | Nd Neodymium 144.24 | Pm Promethium (145) | Sm Samarium 150.36 | Eu Europium 151.96 | Gd Gadolinium 157.25 | Tb Terbium 158.925 | Dy Dysprosium 162.50 | Ho Holmium 164.930 | Er Erbium 167.26 | Tm Thulium 168.934 | Yb Ytterbium 173.05 | Lu Lutetium 174.97 | | | |
| | Ac Actinium (227) | Th Thorium 232.038 | Pa Protactinium 231.036 | U Uranium 238.03 | Np Neptunium (237) | Pu Plutonium (244) | Am Americium (243) | Cm Curium (247) | Bk Berkelium (247) | Cf Californium (251) | Es Einsteinium (252) | Fm Fermium (257) | Md Mendelevium (258) | No Nobelium (259) | Lr Lawrencium (262) | | | |

Metals

- Alkali Metals
- Alkaline Earth Metals
- Lanthanides
- Actinides
- Transition Metals
- Other Metals

Semimetals

- Semimetals

Nonmetals

- Other Nonmetals
- Halogens
- Noble Gases

Key

- Atomic Number
- Symbol
- Name
- Atomic Weight

- **Groups 1-2:** the number of valence electrons equals the element's group number on the periodic table. For example, calcium is in group 2 and has two valence electrons.
- **Groups 13-18:** the number of valence electrons is ten fewer than the element's group number. For example, oxygen is in group 16 and has six valence electrons.
- **Groups 3-12:** the process is not so simple for the transition metals. Transition metal atoms can use electrons from their inner shells as valence electrons, so the number of valence electrons varies.
- **Group 18:** the noble gases in the last group on the periodic table each have eight valence electrons. Since eight is the maximum number of electrons that can fit into each of these atom's outermost energy level, each of the noble gases has a **full octet**. Atoms are most stable when they have eight valence electrons, making the noble gases generally unreactive.

The noble gas helium is an exception with only two valence electrons. Its outermost energy level is only a 1s orbital, which is full with two electrons, making helium stable.

EXAMPLE:

One beryllium (Be) atom has four electrons and an electron configuration of $1s^2 2s^2$. 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 ($2s^2$) are the outermost electrons for beryllium, so they are beryllium's two valence electrons. The two electrons in the second level ($2s$) are the ones involved when a beryllium atom bonds with other atoms.

DRAWING LEWIS SYMBOLS

The Lewis symbol of one atom of an element depends on the element's chemical symbol and the number of valence electrons that element has.

To draw the Lewis symbol of one atom of an element, use the following steps:

1. Write the chemical symbol of the element. The symbol represents the nucleus and all of the electrons not in the valence shell (the core electrons).
2. Determine the number of valence electrons based on the element's periodic table group.
3. For each of the first four valence electrons, draw a single dot on each side of the chemical symbol.
4. After there is one electron dot on each side, each additional electron can be paired with another electron dot until all valence electrons are shown in the structure.

EXAMPLE:

To draw the Lewis symbol of one atom of nitrogen, begin with the chemical symbol "N." Nitrogen is in group 15. To determine the number of valence electrons, we subtract ten from fifteen and find that nitrogen has 5 valence electrons.

Add the first four valence electrons with one dot on each side of the chemical symbol:

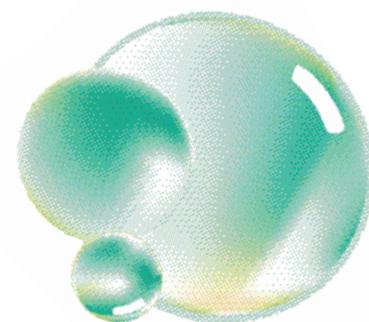


Finally, add one more electron to reach a total of five valence electrons. This is the Lewis symbol for one atom of nitrogen:



Atoms of other elements can be drawn using these same steps, as shown below.

| Atom of Element | Lewis Structure |
|-----------------|-----------------|
| Lithium | Li• |
| Boron | •B• |
| Carbon | •C• |
| Fluorine | :F: |
| Chlorine | :Cl: |



Lewis symbols help illustrate why elements of the same group tend to react similarly. Look at the Lewis symbols for fluorine and chlorine above. The Lewis symbols for elements in the same group have the same number of valence electrons. All of the halogens (group 17) need just one more electron to have a full octet and be stable, so they will all participate in reactions where they gain one electron (see the subsection on [Types of Chemical Bonds](#)).

THE OCTET RULE

The number of valence electrons indicates how many bonds an element is likely to make. The number of electrons that an atom tends to gain, lose, or share (and therefore the number of bonds that atom will make) can be predicted by the **octet rule**, which states that:

- Atoms of main group elements are more stable when they have eight valence electrons, so they *gain, lose, or share electrons by forming chemical bonds with other atoms*. Atoms form bonds in such a way that they achieve an **octet** of valence electrons (eight valence electrons).
- A metal element tends to *lose* electrons until it has the same electron configuration as the noble gas in the period above itself. Sodium tends to lose an electron to achieve the same electron configuration as neon.
- A nonmetal element tends to *gain or share* electrons until it has the same electron configuration as the noble gas in the same period as itself. Sulfur tends to gain two electrons to achieve the same configuration as argon.
- Hydrogen, helium, lithium, and beryllium are exceptions to the octet rule because they achieve the same electron configuration as the noble gas helium, which has only two electrons.

Think About It...

Argon's electron configuration is $1s^22s^22p^63s^23p^6$. What ion of chlorine has the most stable electron configuration? What about calcium?

WRITING CHEMICAL FORMULAS

When atoms form chemical bonds to gain or lose electrons according to the octet rule, those atoms form a compound. Remember, a compound is a pure substance made up of two or more atoms that are chemically combined in a whole-number ratio. The **chemical formula** of a compound shows which elements it contains, as well as how many atoms of each element. 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.

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 symbol for hydrogen

Chemical symbol for oxygen



The **2** indicates that **two** hydrogen atoms are required to make **one** water molecule

When no subscript is written after the chemical symbol, just **one** atom is needed. **One** atom of oxygen is required to make **one** molecule of water.

The subscripts after hydrogen and oxygen in the previous example show how many atoms of each element are needed to make a molecule of water: 2 hydrogen atoms and 1 oxygen atom.



Oxidation numbers are also important when writing chemical formulas. Oxidation numbers are used to show how many electrons an atom has gained, lost, or shared, usually after a chemical reaction. Oxidation numbers are similar to charge. An atom gets a positive oxidation number when it loses electrons, and a negative oxidation number when it gains electrons (just like charge).

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 usually have negative oxidation numbers because they tend to gain electrons.

The overall charge of a compound (or ion) is equal to the sum of the oxidation numbers of all of the elements in that compound. In a neutral compound, the sum of the oxidation numbers is zero. For example, in MgBr_2 magnesium has an oxidation number of +2 and bromine has an oxidation number of -1. There are two bromine ions and one magnesium ion in one formula unit of MgBr_2 , so to find the overall charge:

$$(1 \text{ Mg} \times +2) + (2 \text{ Br} \times -1) = (+2 \text{ from Mg}) + (-2 \text{ from Br}_2) = 0$$

The overall charge is zero, so MgBr_2 is neutral.

Magnesium has a +2 oxidation state because it needs to lose two electrons to achieve an octet, the same electron configuration as the noble gas neon. Bromine has a -1 oxidation state because it gains one electron to achieve an octet, the same configuration as krypton. You can predict what oxidation number an element is likely to have by figuring out how many electrons it can gain or lose in order to achieve an octet.

Quick Fact

Many elements have more than one possible oxidation number. For example, iron may have an oxidation number of +2 or +3 in different compounds. Fe_2O_3 , which contains iron in the +3 state, is the main source of iron used to make steel. FeO , which contains iron in the +2 state, makes up roughly 9% of Earth's mantle. A possible explanation for disturbances in Earth's rotation is that the FeO in the mantle conducts electricity.

Quick Fact

In a chemical formula, the less electronegative element is written first. After reading the subsection on **Electronegativity**, this will be an easier way to determine the order of elements in a chemical formula than oxidation number.

Quick Fact

Chlorine changes to *chloride* in the compound sodium chloride. For some elements, the name is changed to end in *-ide* when the element gains electrons to become an ion. The change reflects a decrease in oxidation number. Because chloride has a negative oxidation number, it is written second in the chemical compound.

EXAMPLE:

Consider writing the formula for sodium chloride, NaCl . Sodium is in group 1 of the periodic table, therefore it needs to give away one electron to complete its octet and to have the same electron configuration as a noble gas (in this case, neon). When sodium reacts with chlorine, sodium gives away one electron, resulting in an oxidation number of +1 for sodium.

Chlorine takes the electron from sodium to complete its octet and obtains an oxidation number of -1. After chlorine gains one electron, it has the same electron configuration as the noble gas argon.

Sodium is written first because its oxidation number is positive. Sodium and chlorine combine in a 1:1 ratio because each atom of sodium wants to give away one electron and each atom of chlorine wants to gain one electron.

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 ₂ H ₅ OH |
| Bleach (chlorine bleach) | Sodium hypochlorite | NaOCl (aq)* |
| Chloroform | Trichloromethane | CHCl ₃ |
| Laughing gas | Nitrous oxide (dinitrogen monoxide) | N ₂ O |
| Lye | Sodium hydroxide | NaOH |
| 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 ₃ COOH (aq)* |
| Water | Water | H ₂ O |
| Wood alcohol | Methanol (methyl alcohol) | CH ₃ OH |

* 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 $\text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$, carbonic acid, H_2CO_3 , is dissolved in water.

** 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 salts, for health reasons.

NAMING CHEMICAL COMPOUNDS

Chemical nomenclature is the system used for naming chemical substances.

There are millions of identified chemical compounds, 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. Similar compounds, like acids or ionic compounds, have similar names.

Quick Fact

There are exceptions for naming compounds. According to the guidelines for naming compounds, the chemical name of water should be dihydrogen monoxide. Instead, the IUPAC name for water is actually oxidane. However, 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 valence electrons. Atoms that have lost electrons and acquired a positive charge are called **cations**. Atoms that have gained electrons and acquired a negative charge are called **anions**. Ions formed from a single atom are called **monatomic ions**.

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

| Symbol | 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 take on different charges 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 | IUPAC Name |
|-----------|---------------|
| Fe^{2+} | Iron(II) ion |
| Fe^{3+} | Iron(III) ion |

Scientists sometimes include Roman numerals even for metals that do not form more than one cation. For example, although Al^{3+} is the most common aluminum ion, it is also theoretically possible for aluminum to have a charge of 1+ or 2+. Scientists can write “aluminum(III) ion” to specify which ion is present.

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 | 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.

Ions containing more than one atom are **polyatomic ions**. When writing the name of an ionic compound, the name of the cation is written first, followed by the name of the anion. The overall charge of a polyatomic ion is the sum of the individual oxidation numbers for each atom in that ion.

- Polyatomic *cations* formed from nonmetal atoms have names that end in “-ium.”

| Symbol | 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 “-ide” at the end of the name (as a 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**. When an element can form two different oxyanions with the same charge, the suffix “-ate” is used for the oxyanion with one more oxygen than the oxyanion with the suffix “-ite.” As shown below, sulfate and sulfite have the same charge, but *sulfate* has one more oxygen atom than *sulfite*.

| 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 at the beginning of the name (as a prefix).

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

Ionic compounds are formed from oppositely charged ions bonded together by electrical forces (remember, opposite charges attract). When writing the name of an ionic compound, the name of the cation is written first, followed by the name of the anion.

Quick Fact

In NH_4^+ , each H has a +1 oxidation state and the N has a -3 oxidation state. To find the overall charge:

$$(1 \text{ N} \times 3-) + (4 \text{ H} \times 1+) = 1+$$

The overall charge is 1+.

NAMING ACIDS

Chemical compounds that release hydrogen cations (H^+) when they are dissolved in water are **acids** (see the [Acids, Bases, and pH](#) section). Many acids are composed of two parts: negatively charged ions (monatomic or polyatomic) combined with enough positively charged hydrogen ions (H^+) to make the compound electrically neutral.

Some acid compounds have only two elements: hydrogen and one other element. These compounds are called *binary acids* and do not contain oxygen.

- To name *binary acids* use the following steps:
 1. Start with the prefix “hydro-.”
 2. Add the name of the monatomic anion (the element that is not hydrogen).
 3. Change the “-ide” ending to “-ic.”
 4. Add the word “acid” to the end of the name.

EXAMPLE:

For the acid HF, begin with the prefix “hydro.” Then, add on the name of the monatomic anion of fluorine, which is fluoride. The result of the first two steps is “hydrofluoride.” Next, change the “-ide” ending to “-ic.” This gives the name “hydrofluoric.” Finally, add the word “acid.” The name of the compound HF is hydrofluoric acid.

| Symbol/Formula | IUPAC Name |
|------------------|--------------------|
| HCl | Hydrochloric acid |
| H ₂ S | Hydrosulfuric acid |

Other acidic compounds contain oxygen. These acids are called *oxyacids*. There are two main methods for naming acids that contain oxygen, depending on the name of the polyatomic anion involved.

- To name *oxyacids that contain an anion ending in “-ate”* use the following steps:
 1. Start with the name of the anion.
 2. Change the “-ate” ending to “-ic.”
 3. Add the word “acid.”

EXAMPLE:

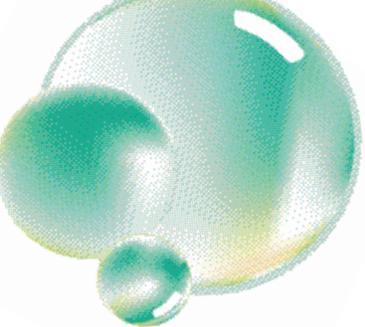
For the acid HClO₃, begin with the name of the anion. ClO₃⁻ is a chlorate ion. Next, change the “-ate” ending to “-ic.” This gives the name “chloric.” Finally, add the word “acid.” The name of the compound HClO₃ is chloric acid.

Quick Fact

Hydrofluoric acid actually refers to HF that is dissolved in water, or HF(aq). When HF is in the gas phase, it is called hydrogen fluoride.

Think About It...

Look back at the definition of an acid. Why do HF(aq) and HF(g) have different names? Is HF(g) still an acid?



| Symbol/Formula | IUPAC Name |
|-------------------------|---------------|
| HNO_3 | Nitric acid |
| H_2SO_4 | Sulfuric acid |

Notice that sulfuric acid is an exception to step 2 above because its name is not “sulfic acid.” Another exception is phosphate, PO_4^{3-} . The acid H_3PO_4 is called phosphoric acid, rather than “phosphic acid.”

- To name *oxyacids that contain an anion ending in “-ite”* use the following steps:
 - Start with the name of the anion.
 - Change the “-ite” ending to “-ous.”
 - Add the word “acid.”

EXAMPLE:

For the acid compound HClO_2 , begin with the name of the anion. ClO_2^- is a chlorite ion. Next, change the “-ite” ending to “-ous.” This gives the name “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 |

Think About It...

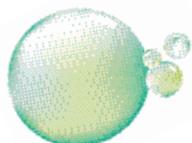
HNO_3 is a neutral compound. The oxidation state of hydrogen is +1 and the oxidation state of oxygen is -2. What is nitrogen's oxidation state? Is this surprising? Why or why not?

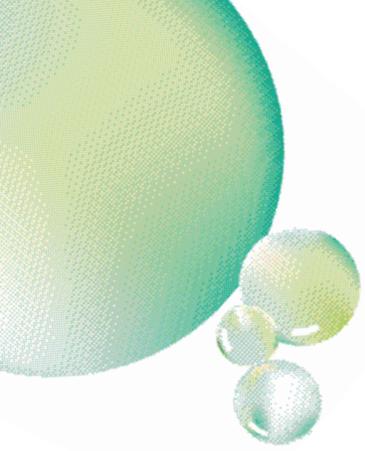


NAMING MOLECULAR COMPOUNDS

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

- To name binary molecular compounds from the molecular formulas use the following steps:
 - Write the name of the element that is the most electropositive (least electronegative). This is usually the element that is farther 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.
 - Write the name of the second element.
 - Change the ending of the second element to “-ide.”
 - Use Greek prefixes on the next page to represent the number of atoms of each element in the compound.





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

Quick Fact
 The prefix “mono—” means one. In carbon monoxide, it indicates that there is only one oxygen atom. The prefix “mono—” is used in many other English words, such as monotonous, which means in one tone or boring.

EXAMPLE:

For the molecular compound **CO**, oxygen has a negative oxidation number because it is found farther to the right on the periodic table, and therefore attracts electrons more strongly than carbon does. Since carbon is found farther to the left on the periodic table, it has a positive oxidation number and is less electronegative. Therefore, carbon is written first. Oxygen is written second, and its ending is changed to “-ide.” At this point, the name is “carbon oxide.” Finally, add the Greek prefix “mono-” to oxygen. The name of the compound is carbon monoxide.

Likewise, the name of the molecular compound **CO₂** is carbon dioxide.

Quick Fact
 The 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. It is also sometimes left off of the second element. For example, HF is hydrogen fluoride, not monohydrogen monofluoride.

| Symbol/Formula | IUPAC Name |
|-------------------|----------------------|
| Cl ₂ O | Dichlorine monoxide |
| NF ₃ | Nitrogen trifluoride |



Remember, these rules apply only to molecular compounds, which are made of nonmetal atoms bonded together. The rules for naming combinations of metal and nonmetal atoms (see the subsection on [Ionic Compounds](#)) are not the same.

NOTES

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 SO₃. The formula for a sulfite ion is SO₃²⁻. One unit of a molecular (covalent) compound such as SO₃ is called a “molecule.” One unit of a polyatomic ion such as SO₃²⁻ is called a “formula unit.”

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 gravity and pressure, a force is any kind of push or pull on an object. Pressure is a type of force that requires objects to touch, so it is known as a *contact force*. 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*. The most familiar examples of non-contact forces are gravity and magnetism.

Although the forces mentioned above are easily identified and experienced, there are other types of forces that play an important role inside chemical substances. These forces include intermolecular and intramolecular forces. An **intermolecular force** is a force acting *between* two or more molecules. An **intramolecular force** is a force acting *within* an atom, ion, or molecule.

INTRAMOLECULAR FORCES

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

1. 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.
2. The charges in an atom push each other apart if they are the same (both positive or both negative).

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

| | |
|---|--|
|  | The electron (-) is attracted to the proton (+) at the center of an atom.* |
|  | The two electrons repel each other, and their like charges push them away from each other.* |
|  | Both of the two electrons are attracted to the proton but still repel one another. They tend to be close to the proton but not close to each other.* |

*In the atom images above, the electron cloud has been removed for easier visualization. The images are not to scale.

These forces also occur in ions and ionic compounds. Just like protons and electrons, cations and anions are attracted to one another. For example, an ammonium cation (NH_4^+) will be attracted to a chloride anion (Cl^-). A bromide anion (Br^-) will be attracted to a potassium cation (K^+).

Coulomb's law and other principles of chemistry, such as quantum mechanics, combine to explain the structure of atoms.

Quick Fact

Quantum mechanics explains why the electrons in an atom stay a certain distance away from the nucleus. Otherwise, protons and electrons would simply crash into one another because of Coulomb's law and the attraction of their charges.

PERIODIC TRENDS

As mentioned previously, **valence electrons** are the electrons in the outermost energy level of an atom (including all sublevels of the outermost energy level). They are represented as the dots that surround the chemical symbol in a Lewis symbol. For group 1-2 elements, the number of valence electrons equals their group number. For group 13-18 elements, the number of valence electrons is ten fewer than their group number (see the subsection on [Lewis Symbols](#)).

- An atom of a main group element can typically hold a maximum of eight valence electrons. The exceptions are hydrogen and helium.
- Atoms that have fewer than eight valence electrons tend to form bonds with other atoms. Atoms will give, take, or share electrons to achieve a full outermost energy level (with eight electrons), which makes them stable.
- Although some energy levels can hold more than eight electrons, main group elements usually 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 ($1s^2 2s^2 2p^6 3s^2 3p^5$) will only use 7 valence electrons in a chemical bond (see the section on Electron Configuration from *The Core of Chemistry*). The third sublevel with energy level three, 3d, is actually higher in energy than the 4s sublevel and therefore will not be filled until after the 4s sublevel is filled.
- Generally, metals lose valence electrons to become more stable and nonmetals gain electrons.

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 its valence electrons are on the fourth energy level. The energy level of potassium's valence electrons can also be seen in its electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$.

ELECTRONEGATIVITY

Electronegativity is a chemical property that describes how strongly the nucleus of an atom attracts the electrons in a chemical bond. Electronegativity indicates how strong of a pull an atom has on electrons and how strongly it will compete for electrons during reactions. When two atoms are bonded together, the electrons are more attracted to and more tightly bound to the more electronegative atom. The difference in electronegativity between two atoms determines what type of bond they form.

Quick Fact

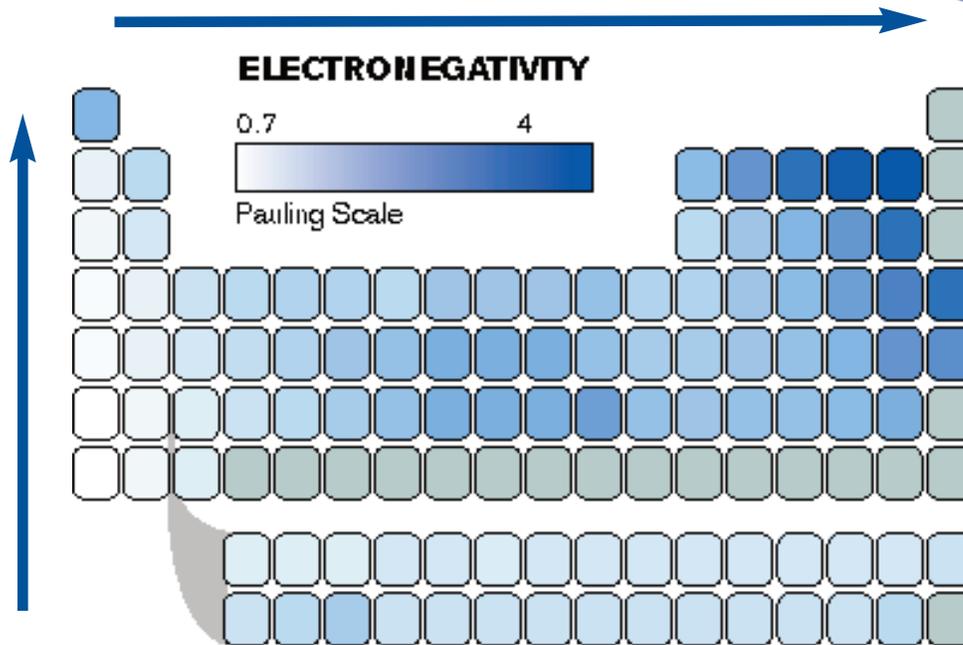
The first energy level of an atom can hold a maximum of two electrons. Hydrogen and helium are the only elements with electrons that **only** occupy this energy level. The electrons of all other elements occupy additional energy levels.

- **Within a period, electronegativity tends to *increase* from left to right.** Within the same period of the periodic table, moving one column to the right means adding one more proton and electron to the same outer electron shell (elements within the same period have the same number of electron shells). Think of proton-electron attractions as small magnets. The more proton-electron pairs are present in an atom, the more the outer shell is drawn towards the center of the atom. As the outer shell is drawn in and the valence electrons get closer to the attractive, positively charged nucleus, the atom holds onto its own electrons and other electrons more tightly.
- **Within a group, electronegativity tends to *decrease* from top to bottom.** Moving down one group on the periodic table means adding one more electron shell, increasing the size of the atoms. As the atoms get bigger, there is more space and there are more electrons between the nucleus and the outer shell. The nucleus' ability to attract an electron decreases as the electron moves farther away from the nucleus and as more electron shells are added between that electron and the nucleus—this is called “shielding”. Therefore, the nucleus of a larger atom can't attract other (or its own) electrons as well as a smaller atom.

Fluorine and chlorine have the same number of valence electrons, but fluorine is smaller, therefore fluorine is more capable of attracting electrons and adding them to its shell to complete its octet. Fluorine is more electronegative than chlorine (See the subsection on [Atomic Radii](#)).

Think About It...

How does electronegativity explain why metals lose electrons and nonmetals gain electrons in chemical reactions? Go back to the example on page 8. How does electronegativity explain the reaction between sodium and chlorine?



*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. This includes all the noble gases besides krypton and xenon, as well as elements 103 through 118.
 ***The arrows indicate the *general trend* of electronegativity: it increases moving to the right within a period and moving up within a group.

The most strongly 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, meaning its nucleus attracts electrons more weakly than any other element. Francium is the most “electropositive” element.

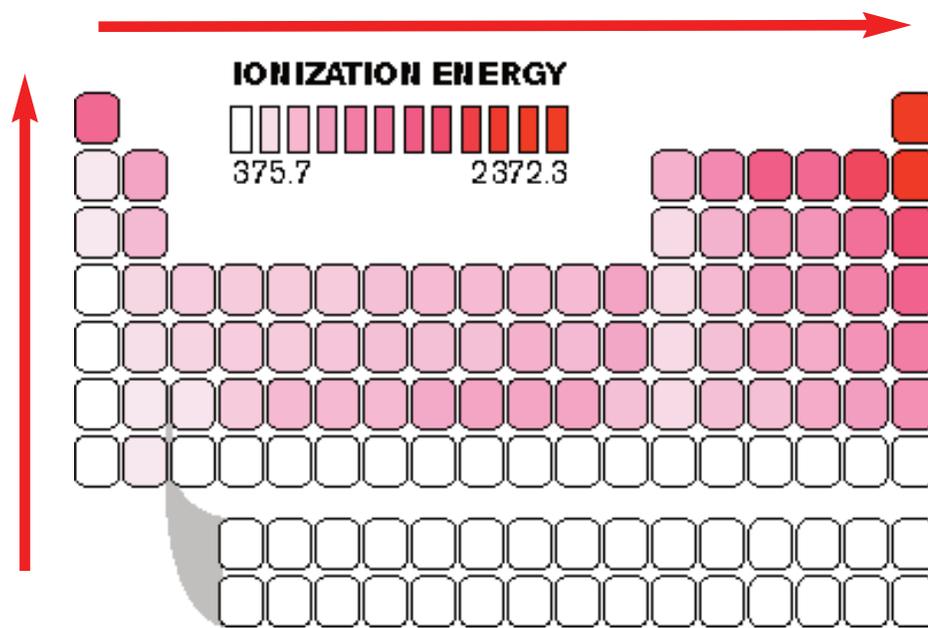
EXAMPLE:

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's outermost energy level, level 2, already has eight valence electrons (an octet). Since the outer shell is full, neon does not need to gain or lose electrons - it is already stable!

IONIZATION ENERGY

Ionization energy is the amount of energy needed to remove one of the outermost electrons from a neutral atom in the gas phase. Each negatively charged electron is attracted to the positively charged nucleus, so in order to remove an electron a certain amount of energy (the ionization energy) is needed to overcome that attraction. Think of ionization energy as an indicator of how strongly an atom holds onto its valence electrons. A higher ionization energy means the atom has a stronger hold on its valence electrons, so it takes more energy to get the electrons away from the atom and its nucleus. A lower ionization energy means that the atom can more easily lose its valence electrons, and the electrons are not as tightly bound to the nucleus.

- **Within a period, ionization energy tends to *increase* from left to right across the periodic table.** Much like electronegativity, this trend occurs because of the increasing number of protons and electrons. More protons and electrons means more proton-electron pairs attracted to one another (again, think of proton-electron attractions as small magnets). As proton-electron pairs are added, the electron shell is pulled closer to the nucleus. Electrons closer to the nucleus are more strongly attracted to the nucleus, so it requires more energy to overcome that attraction and remove an electron.
- **Within a group, ionization energy tends to *decrease* from top to bottom down the periodic table.** Also similar to electronegativity, this trend occurs because atoms get larger moving down a group. The valence electrons are in electron shells farther and farther from the nucleus. Less energy is needed to remove an electron that is farther from the nucleus, so ionization energy decreases as atoms get larger moving down a group.



*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 moving to the right within a period and moving up within a group.

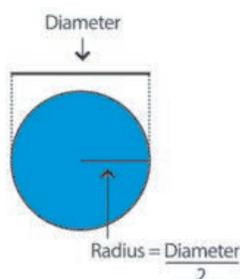
Ionization energy and electronegativity trends explain why metals and nonmetals react differently to achieve stability. Metallic elements like sodium lose, rather than gain, electrons to reach an octet because they have lower ionization energies (and lower electronegativity). Moving to the right across the periodic table, atoms get closer to having a full octet, become more electronegative, and have higher ionization energy. As a result, nonmetals are more likely to gain, rather than lose, electrons and become like the noble gas in their same period, rather than losing electrons to become like the noble gas in the period above.



Since atoms want to be stable, elements like the halogens (fluorine, chlorine, etc.) that are only one electron away from a full octet will want to gain an electron, and it will take more energy to take one away. Notice that noble gases have higher ionization energies because they already have an 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. The radius of the ball 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.

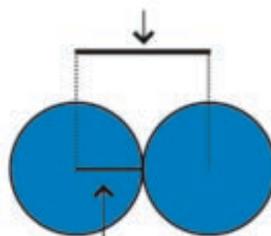


An atom does not have a clearly defined edge because of the electron cloud. The atomic radius is determined by how close one atom is to a neighboring atom. It is half the distance between the nuclei of two atoms of one element that are bonded together. This distance is so small that it is typically measured in picometers ($1 \text{ pm} = 10^{-12} \text{ m}$).

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.

Distance between the nuclei of two neighboring sodium atoms:
372 pm



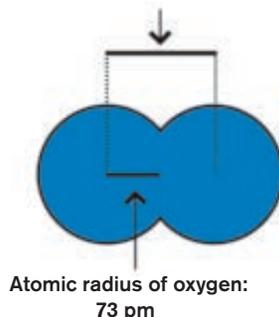
Atomic radius of sodium:
186 pm

EXAMPLE:

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



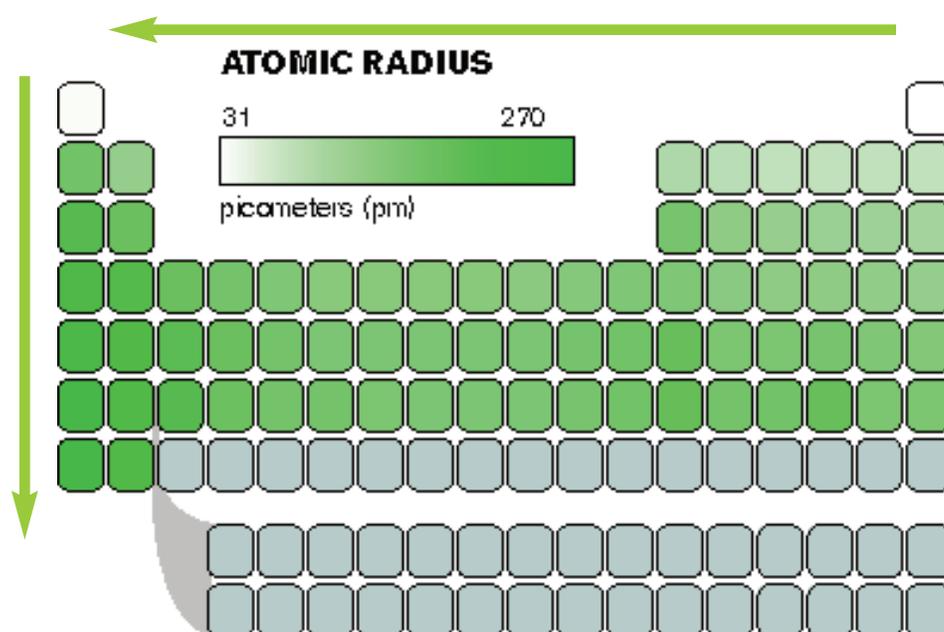
Distance between the nuclei of oxygen atoms joined in a diatomic oxygen molecule:
146 pm



Quick Fact

The cation of an element generally has a smaller radius than the neutral atom. This is because electrons are attracted to the positive protons in the nucleus, but are repulsed by other negatively charged electrons. The electrons want to be near the nucleus but far away from each other. Removing an electron decreases the electron-electron repulsion, so the nucleus can pull the remaining electrons in more tightly.

- **Atomic radius tends to *decrease* from left to right across the periodic table.** Similar to electronegativity and ionization energy, this trend results from the increasing positive charge in the nucleus. The more positively charged nucleus attracts the electrons more strongly and pulls them closer to the nucleus.
- **Atomic radius tends to *increase* from top to bottom down the periodic table.** This trend is caused by the increasing energy levels (and thus larger orbitals). The electrons are farther from the pull of the positive nucleus, and are therefore not held as tightly or as closely.



*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 the atomic radii of elements can help to explain their ionization energies. If the radius of an element is small, the electrons are closer to the nucleus. This means that the protons in the nucleus are more strongly attracting and pulling in the electrons in the outer shells. The amount of energy needed to pull an electron off (the ionization energy) gets smaller as the distance between the valence electrons and nucleus gets larger, and as more electrons are added between the valence electrons and nucleus. When there are more electrons shells, the attraction between the nucleus and the outermost electrons is weaker, so pulling off an electron will be easier and require less energy.

TYPES OF CHEMICAL BONDS

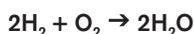
CHEMICAL BONDS

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. Ionic bonds, for example, are the result of attraction between oppositely charged 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 (H₂). Likewise, a molecule of oxygen gas (O₂) contains two bonded oxygen atoms. When a molecule of oxygen gas combines with two molecules of hydrogen gas, the hydrogen-hydrogen bonds in H₂ and the oxygen-oxygen bonds in O₂ break. Each oxygen atom then forms two new bonds, each to a different hydrogen atom, producing H₂O molecules (see the **Chemical Reactions section).**



There are three types of chemical bonding: ionic, covalent, and metallic. Bonding involves only the valence electrons of an atom (see the subsection on **Valence Electrons**).

IONIC BONDING

Ionic bonds occur when one atom gives up electrons and another atom takes them. The atom that gains electrons becomes a negative ion (anion). The atom that loses electrons becomes a positive ion (cation). The resulting ions have opposite charges and become attracted to one another. This force of attraction holds the ions together. Compounds held together by ionic bonds are called **ionic compounds**.

Ionic bonds typically occur between metal atoms and nonmetal atoms because of their very different electronegativities. The less electronegative atom (the metal) readily gives up electrons, and the more electronegative atom (the nonmetal) readily accepts these electrons. The resulting metal cation and nonmetal anion are attracted to each other because of their opposite charges. The cations and anions in an ionic bond can also be polyatomic.

As a result, individual atoms or ions are not held together closely by shared electrons, as is the case with covalent bonds. There are no distinct, individual molecules, but rather a collection of ions held together by the attraction of their charges. Ionic compounds exist as three-dimensional networks of ions all connected by ionic bonds.

EXAMPLE:

When sodium (Na) and chlorine (Cl) combine to make sodium chloride (NaCl), the chlorine atoms want to take the valence electrons from the sodium atoms. Chlorine, a nonmetal, is on the more electronegative side of the periodic table. Sodium, a metal, is on the electropositive side and donates electrons to the chlorine atoms.

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

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

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

Notice that the electron produced in Step 1 is used in Step 2, so it is cancelled out in the overall 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: $\text{Mg} \rightarrow \text{Mg}^{2+} + 2 \text{ electrons}$ (production of an Mg^{2+} cation plus release of electrons)

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

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

Magnesium is in group 2 and has two valence electrons. When magnesium reacts with chlorine, each Mg atom donates two electrons. Chlorine is in group 17 and has seven valence electrons, so each Cl atom will only gain one electron. Therefore, two Cl atoms are needed to accept the electrons from each Mg atom, making the 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 bond with one magnesium atom.

Here's the periodic table 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 nonmetal atoms, which have similar electronegativities and are close to one another on the periodic table. Compounds formed entirely from atoms that share electrons through covalent bonds are called **covalent compounds**.

- Covalent bonds create stable compounds if the sharing of electrons brings about a noble gas electron configuration for each atom (with eight valence electrons).
- In a covalent bond, one atom does not actually lose an electron in order for the other atom to gain an electron. Instead, the atoms **share** the electrons. Both atoms in a covalent bond gain electrons. Two atoms would not be able to share electrons if their electronegativities were very different, because one atom would attract the electrons much more strongly than the other.
- Each covalent bond contains one electron pair (two electrons).
- Nonmetal atoms generally have high electronegativity, meaning that they want to gain electrons, and also high ionization energy, meaning that they do not want to give up their own electrons. Nonmetal atoms tend to form covalent bonds because it allows them to share electrons, so they can gain electrons without giving up any of their own.

Quick Fact

Hydrogen is an exception to the octet rule. A hydrogen atom is stable when it makes one bond (has two valence electrons), because it achieves the same electron configuration as the noble gas helium.



EXAMPLE:

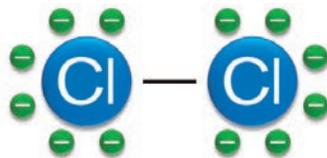
Look at the molecule of Cl_2 below. Chlorine is found on the third period of the periodic table and is part of the halogen group. Because this molecule is made up of two atoms that are the same, both atoms in the bond have the same electronegativity and the same ability to attract electrons. The two chlorine atoms share the electrons equally between them. Remember that only the outermost electrons participate in the bond. Chlorine's electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^5$, so its outermost electrons are those in the *third* energy level: the $3s^2 3p^5$ subshells. Chlorine has a total of seven valence electrons, and it is these electrons that 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. Both atoms now have the same electron configuration as the noble gas in the same period, argon, and are therefore stable.

A *single covalent bond* contains two shared 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 between the bonded atoms. Cl_2 is shown as:



- The bond that forms between the chlorine molecules to make Cl_2 is called a *single covalent bond*. Chemical compounds can also contain *double covalent bonds* and *triple covalent bonds*.

- **Single covalent bond (single bond):** a covalent bond sharing only one pair of electrons (two electrons total) between two atoms.
- **Double covalent bond (double bond):** a covalent bond sharing two pairs of electrons (four electrons total) between two atoms.
- **Triple covalent bond (triple bond):** a covalent bond sharing three pairs of electrons (six electrons total) between two atoms.

Atoms do not have to be identical to form a covalent bond. A covalent bond can form between two nonmetals or between a nonmetal and a semimetal.

Quick Fact

Structures that show atomic centers (symbols) and either lines or dots for the valence electrons, such as the one pictured to the left, are called Lewis structures (see the section on [Lewis Structures](#)).

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

Think About It...

What type of covalent bond is found in diatomic oxygen? What about diatomic nitrogen? Use the periodic table to determine how many bonds oxygen and nitrogen tend to make.

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 metal atoms bond by contributing their electrons to a “sea” of shared electrons. This “sea” of electrons is shared among all of the bonded atoms in 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. Each metal atom gives up its valence electrons, forming a “sea” of electrons.
 - The creation of an electron “sea” only occurs if there are no elements with high electronegativity (i.e. nonmetals) present to accept the electrons. Since there are no nonmetals to accept the electrons, there is no more stable place for the electrons to be than in the electrons sea, shared between all metal atoms.
 - Remember, metals often have low electronegativity and low ionization energy. This allows metal atoms to give up electrons easily. Metallic bonds occur when the Coulombic attraction keeping the electrons bound to the atom’s nucleus is weaker than the electron’s energy; the electrons have enough energy to break free.
- Metallic bonds can also form among elements that have higher 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. Atoms of gold will not give up electrons to other substances, but will come together to form strong metallic bonds.
- Some metallic elements are actually quite easy to keep in pure form because they are relatively unreactive. These elements include gold, platinum, and silver.

Since electrons in a metallic compound are not held tightly to individual nuclei, the electrons are able to move more freely. Many properties of metals result from the high mobility of electrons in a metallic bond, as well as the ability of those electrons to move across the entire object.

- **Luster:** the ability to reflect light. Luster is why metals look shiny.
 - 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:** a measure of the rate at which thermal energy can travel through a material.
 - Metals also have good thermal conductivity. As part of a metal is heated, the electrons become excited. The excited electrons then travel to the other side of the metal, carrying the thermal energy with them. Electrons are much better at carrying this energy than nuclei, so metals conduct thermal energy better than substances whose electrons cannot move freely. This is why cold metal feels colder to the touch than equally cold plastic or wood. The thermal energy in a person’s hand is transferred to metal more quickly than it is transferred to wood, even if the difference in thermal energy (difference in temperature) between someone’s hand and the wood is the same.

Quick Fact

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

Think About It...

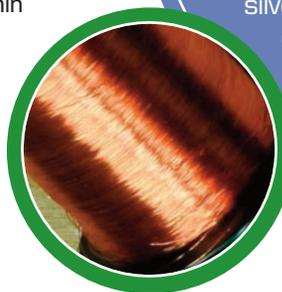
In the movie *A Christmas Story*, Flick is dared to touch his tongue to a metal flagpole in freezing temperatures. He takes the dare, and his tongue gets stuck to the metal. Why would this NOT have happened if the flagpole had been made of wood?

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

| | |
|--------------------------------|--|
| <p>Ionic Bonding</p> | |
| <p>Covalent Bonding</p> | |
| <p>Metallic Bonding</p> | |

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

LEWIS STRUCTURES

Lewis structures are one way to represent chemical compounds. In the Lewis structure of a compound, chemical symbols are used to represent each atom in the compound. Bonds (lines) are drawn between symbols to show how electrons are shared. Dots represent valence electrons that are not shared, or “lone pairs” belonging to only one atom.

When atoms react, they create bonds to achieve a full octet of eight valence electrons (again, hydrogen and helium are exceptions—they are stable when they have two valence electrons). The **bonding rules** below determine the number of bonds that one atom of an element will make in a compound.

- **Gaining electrons:** when an atom is gaining electrons, the number of bonds it will make equals the number of electrons that the atom needs to complete its octet. In a covalent bond, both atoms gain electrons. In an ionic bond, the more electronegative atom gains electrons.

Number of bonds = 8 - Number of valence electrons

For example, fluorine is in group 17 and has seven valence electrons. According to the equation above, fluorine makes one bond because $8 - 7$ valence electrons = 1 bond. Moving along period 2, oxygen makes two bonds, nitrogen makes three bonds, and carbon makes four bonds. Nonmetals tend to gain electrons.

- **Losing electrons:** when an atom is losing electrons, the number of bonds it will make equals the number of valence electrons that atom has. In ionic compounds, the less electronegative element loses all of its high-energy valence electrons to achieve the same electron configuration as the noble gas in the period above it.

Number of bonds = Number of valence electrons

In period 2 of the periodic table, lithium makes one bond and beryllium makes two bonds. Metals tend to lose electrons.

- Hydrogen atoms always make one bond, whether they are in an ionic or a covalent compound. Hydrogen atoms are never given a full octet because a hydrogen atom can hold only two electrons.
- Except in rare and unstable cases, carbon cannot have any lone pairs and will always make four bonds.

COVALENT COMPOUNDS

Covalent compounds are made up of one or more nonmetal atoms connected by covalent bonds. In the Lewis structure of a covalent compound, lines between atoms are used to represent the sharing of electrons (covalent bonds).



Think About It...

How many bonds is krypton likely to make? How does krypton's electron configuration explain this number?

Think About It...

Why can't a hydrogen atom hold more than two electrons? What is the ground-state electron configuration of a hydrogen atom?

At normal room temperature, halogens like bromine exist as diatomic molecules. The Lewis structure of diatomic bromine, Br₂, a covalent compound, is shown below:



- The Lewis symbols of the individual bromine atoms are shown to the left of the arrow.
- The Lewis structure of the covalent compound Br₂ is shown to the right of the arrow.
- All of the black dots around each of the bromine atoms represent lone electron pairs belonging to that atom (each dot is one electron). The lines between atoms are bonds (shared electron pairs).
- The electron pair shared between the bromine atoms in Br₂ is a single covalent bond. The bond is made up of one electron from each bromine atom. Both electrons in the bond are shared equally, so that each bromine atom gains one electron and achieves an octet (eight valence electrons).

To count the number of valence electrons that one atom has in the Lewis structure of a covalent compound:

$$\text{Valence electrons} = \left(\begin{array}{c} \text{Number of electron dots} \\ \text{around the atom} \end{array} \right) + \left(\begin{array}{c} 2 \times \text{Number of bonds the atom} \\ \text{makes in the compound} \end{array} \right)$$

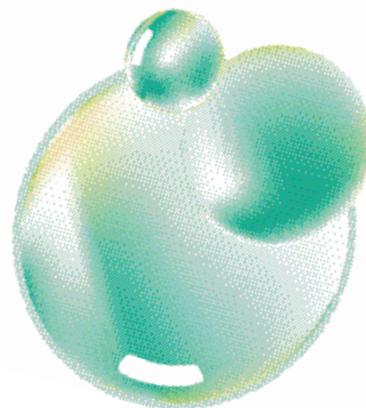
Each bromine atom in the Lewis structure of Br₂ has (6 electron dots) + (2 x 1 bond) = 8 valence electrons.

To draw the Lewis structure of a covalent compound:

1. Count the total number of valence electrons in the compound. This is the sum of the valence electrons of all of the atoms in the compound.
2. Represent each atom using its chemical symbol. Draw bonds between atoms according to the bonding rules on the previous page.
3. Count the number of valence electrons that each atom has after adding the bonds in step 2. Give each atom (besides hydrogen) a complete octet by adding electron dots until the atom has eight valence electrons.
4. Count the total number of valence electrons in the Lewis structure:

$$\text{Total valence electrons} = \text{Total number of electron dots} + 2 \times \text{Total number of bonds}$$

The total number of valence electrons in the Lewis structure should be the same as the total valence electrons from step 1.



EXAMPLE:

To draw the Lewis structure of one molecule of water, begin by counting the total number of valence electrons in H₂O.

$$\begin{array}{r} (1 \text{ valence electron per H atom}) \times (2 \text{ H atoms}) = 2 \text{ valence electrons from hydrogen} \\ (6 \text{ valence electrons per O atom}) \times (1 \text{ O atom}) = 6 \text{ valence electrons from oxygen} \\ \hline \text{One H}_2\text{O molecule: } 8 \text{ valence electrons} \end{array}$$

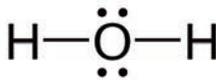
One molecule of H₂O has a total of 8 valence electrons.

Next, draw bonds between the chemical symbols of hydrogen (H) and oxygen (O). Let's look at the bonding rules. Oxygen is in group 16, so it should make 8 – 6 valence electrons = 2 bonds. Hydrogen makes 1 bond. This gives us:



Each oxygen atom now has 2 x 2 bonds = 4 valence electrons.

Oxygen needs 4 more electrons to complete its octet, so we add 4 electron dots around oxygen:



Remember, hydrogen does not need a full octet and is stable with only two valence electrons.

Finally, count the total number of valence electrons in the Lewis structure.

$$\text{Total valence electrons} = (4 \text{ electron dots}) + (2 \text{ electrons per bond}) \times (2 \text{ bonds})$$

$$\text{Total valence electrons} = 8 \text{ valence electrons}$$

This is the same as the number of valence electrons we counted at the beginning! The image above is the Lewis structure for one molecule of H₂O.

IONIC COMPOUNDS

Ionic compounds are made up of two or more ions held together by ionic bonds. Unlike in covalent compounds, ionic compounds are not held together by shared electrons. Instead, ionic compounds are held together by the attraction between opposite charges (cations and anions). Because ionic compounds are a collection of ions, and not atoms held together by individual bonds, Lewis structures are not used as frequently to draw ionic compounds.

The Lewis structure for an ionic compound is simply all of the individual Lewis symbols for each of the ions in the compound, as shown on the following page.

Think About It...

What does the Lewis structure of the covalent compound NH₃ (ammonia) look like? How many bonds does one nitrogen atom tend to make?

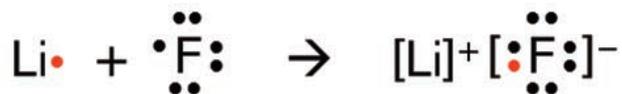
Quick Fact

One unit of a covalent compound like SO₃ is called a "molecule," but one unit of an ionic compound like NaCl is called a "formula unit." This is because ionic compounds do not have discrete (separate) molecules the way that molecular compounds do. Ionic compounds are a collection of ions held together by the attraction between opposite charges, rather than individual atoms bonded by shared electrons.

To write the Lewis symbol of an ion from the chemical formula:

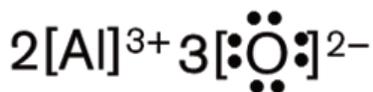
1. Start with the Lewis symbol of the neutral atom (see the subsection on [Lewis Symbols](#)).
2. Add or remove electron dots to show how many electrons the neutral atom has lost or gained.
3. Put the symbol in brackets, and add the charge as a superscript outside of the brackets.

Lithium fluoride (LiF) is an example of an ionic compound:



- The neutral atoms of lithium and fluorine are shown to the left of the arrow. The Lewis structure of the ionic compound LiF is shown to the right of the arrow.
- Lithium, the less electronegative atom, donates one electron to fluorine, the more electronegative atom.
- The lithium cation and the fluorine anion are attracted to each other because of their opposite charges.
- Both lithium and fluorine reach a full outer energy level and the same electron configuration as a noble gas. Lithium's electron configuration goes from $1s^2 2s^1$ as a neutral atom to $1s^2$ as a $1+$ cation, the same configuration as helium. Fluorine's electron configuration goes from $1s^2 2s^2 2p^5$ as a neutral atom to $1s^2 2s^2 2p^6$ as a $1-$ anion, the same configuration as neon.

Many ionic compounds are not formed in a 1:1 ratio, so coefficients can be used to show how many ions of each element are present. For example, aluminum and oxygen combine in a 2:3 ratio to form aluminum oxide, Al_2O_3 . Aluminum oxide can be represented as shown below:



Quick Fact

Both rubies and sapphires are made of a mineral form of aluminum oxide called corundum. When small amounts of chromium are present in corundum, the mineral appears red and is called ruby. When small amounts of iron and titanium are present, the mineral appears blue and is called sapphire.



Think About It...

Why do aluminum and oxygen combine in a 2:3 ratio? In what ratio would aluminum and fluorine combine?

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 orbiting around the sun. On the earth, all objects are pulled toward the earth's center. Think of a book that slides over the edge of a table or a backpack being dropped. The book and the backpack drop to the ground because of gravity.

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

One factor that affects the gravitational force (gravitational attraction) between two objects is the mass of each object. Objects with greater mass have greater gravitational force. Therefore, a backpack will fall toward the earth because the mass of the earth is much, much greater than the mass of someone's 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 between them. 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.

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 section in *The Core of Chemistry*).

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:

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

For any two objects, F_g is the gravitational force between them, 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 Newton's 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 his 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

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. 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). As a result, metals with magnetic properties are attracted to other magnets. 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.

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

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

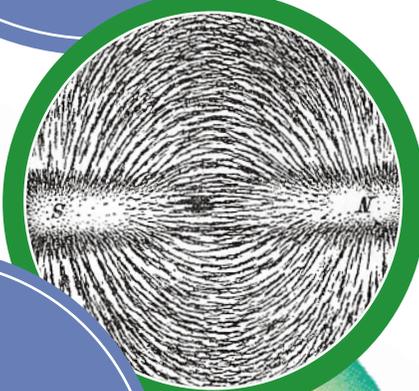
In addition, some magnets have a stronger magnetic force than others. There are two basic kinds of magnets—permanent and temporary. A **permanent magnet** keeps a certain level of magnetism for a long time. A **temporary magnet** acts like a permanent magnet when it is within a strong magnetic field.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

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.

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.



Quick Fact

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

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.

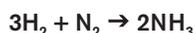
- **Reactants:** the starting material or materials for a chemical reaction.
- **Products:** the substance or substances produced from a chemical reaction. Sometimes one or more of the products can be classified as byproducts.

In general, a chemical reaction will be represented as shown below:

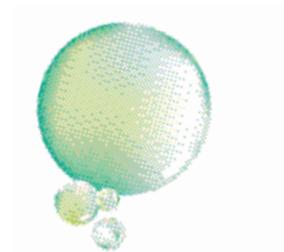
Reactants → Products

EXAMPLE:

The chemical reaction between hydrogen and nitrogen is shown below:



- The hydrogen, H_2 , and nitrogen, N_2 , molecules are the reactants.
- The resulting ammonia, NH_3 , is the product.



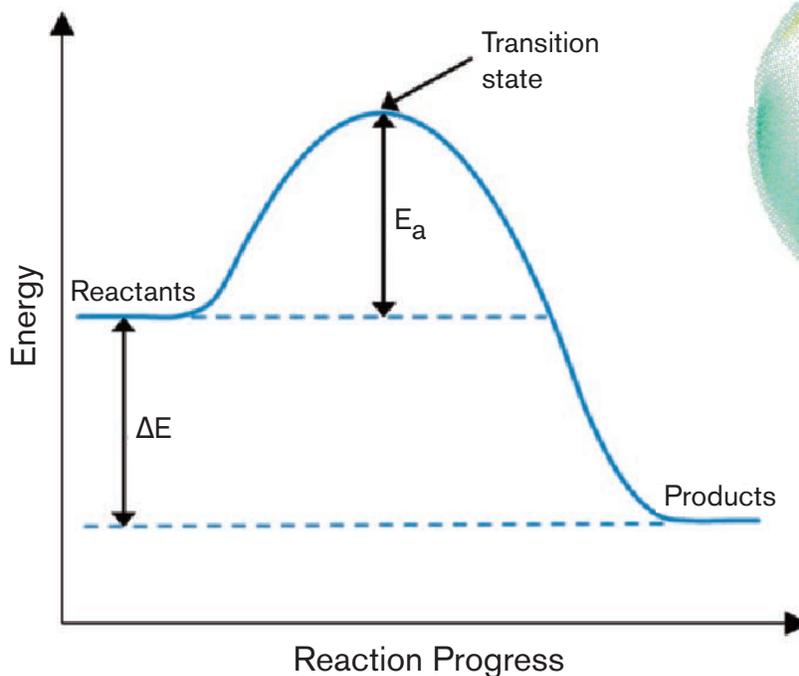
VISUALIZING CHEMICAL REACTIONS

When a chemical reaction occurs, the reactants do not just change directly into the products. It takes time and energy for chemical bonds to be broken and made, and the reactants are often rearranged in multiple different ways as the atoms shift around to become the final product. The road that atoms and molecules take to get from starting materials to products is called the **reaction pathway**.

As the reactants change and bonds are broken and remade, the total energy of the reactants changes too. A **reaction coordinate diagram** plots the energy of the reaction—and all of the molecules and atoms involved—versus how far the reaction has progressed, as shown on the following page. When all of the molecules and atoms are lower in energy, they are more stable. Conversely, they are less stable when they are higher in energy. In general, chemical compounds try to move towards a state that is lower in energy because it is more stable.

Think of a reaction pathway as a hiking trail going over a mountain, and think of altitude (height) as the total energy of the reactants. The altitude changes moving along the trail, and a hiker has to pass through all of the different altitudes along the trail to get from the beginning of the trail to the end. Similarly, reactants have to progress along the reaction pathway, changing energy as they go, until they reach the final products.

Reaction Coordinate Diagram



- The initial flat portion is the energy of the reactants. The flat portion at the end is the energy of the products.
- ΔE represents the overall change in energy. It is the difference between the starting energy and the final energy.

$$\Delta E = (\text{Energy of the products}) - (\text{Energy of the reactants})$$

- E_a represents the **activation energy**. Activation energy is the minimum amount of energy needed to start a chemical reaction.
 - The reactants pass through a state that is much higher in energy in order to get to the products, called the “transition state.” In the transition state, the bonds in the reactants are in the process of breaking and the bonds in the products are in the process of forming. The transition state is the top of the hill (highest in energy) in the reaction coordinate diagram.
 - Activation energy is the amount of energy required to get the reaction “over the hill,” or to start the reaction by overcoming the energy barrier. It is the energy needed to go from the reactants to the transition state.

$$E_a = (\text{Energy of the transition state}) - (\text{Energy of the reactants})$$

- All chemical reactions need some amount of activation energy. The activation energy is needed to break existing bonds so that new bonds can be made.

Quick Fact

A spark provides the energy needed to start a combustion reaction and set wood on fire. The thermal energy (heat) of the spark is greater than the activation energy of the combustion reaction.

The activation energy (E_a) and the change in energy (ΔE) above are independent of one another. The change in energy depends only on the starting energy and the final energy (the reactants and the products) and is not affected by energies along the way.

RATES OF CHEMICAL REACTIONS

Different reactions occur at different rates (or speeds), and even the same reaction can occur at different rates under different conditions. **Chemical kinetics** is the field of science that studies the rates of chemical reactions. Kinetics answers the question: “how fast will this reaction occur?”

To understand reaction rates, it is helpful to first understand what actually happens during a chemical reaction. Three things must be true for a chemical reaction to occur:

1. Particles of the reactants must collide with (or run into) each other. “Particle” here just means one unit, so a particle of a reactant could be an atom, an ion, a molecule, or a formula unit.
2. The particles must have enough energy to react—this usually means the particles must be going at least a certain speed. “Enough energy” means at least the activation energy. If the particles are not moving fast enough, they will not have enough energy to break old bonds and form new bonds when they collide.
3. The particles must have the proper orientation to react.

Particles are always moving around randomly, so sometimes they will collide with one another. If the particles collide with enough energy and the right orientation, they react. The frequency of particles colliding with each other successfully and reacting determines the reaction rate. This is called **collision theory**.

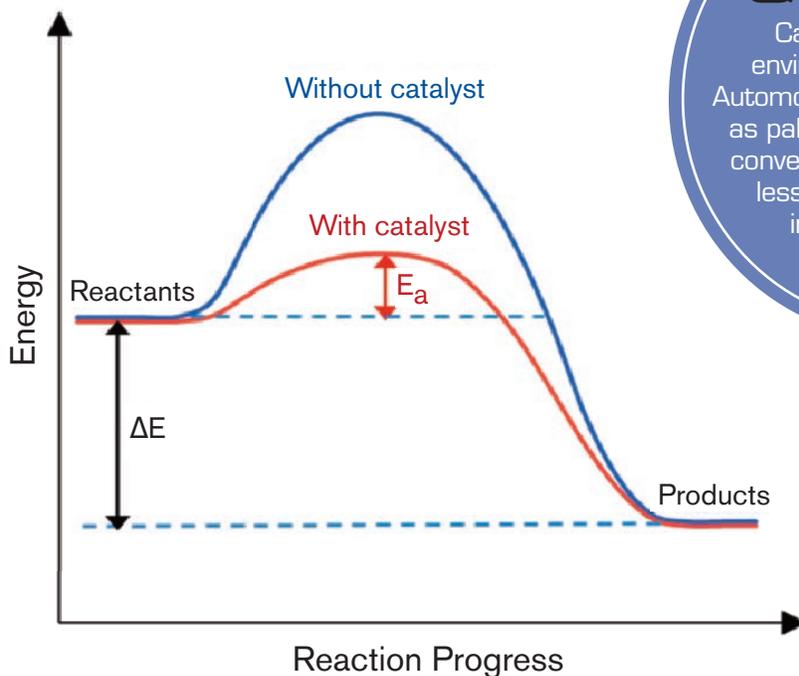
In general, the rate of a chemical reaction can be increased in a few different ways:

1. **Concentration:** A higher concentration of the reactants increases the number of particles in the same amount of space, so it is more likely that particles will run into each other. Think of particles as people. People are more likely to run into each other in a crowded room. Generally, increasing the concentration of the reactants increases the reaction rate.
2. **Temperature:** Remember, temperature is a measure of kinetic energy, or how fast the particles in a substance are moving. When particles are moving faster, they are more likely to collide. People running around in a room are more likely to run into each other than people walking. The particles also have more energy (on average), so when two particles collide they are more likely to have enough energy to react.
3. **Catalysts:** A catalyst is a substance that changes the rate of a reaction. Catalysts are most often used to increase the reaction rate (to make the reaction happen more quickly) by making it easier for the reaction to take place. The identity and amount of a catalyst is the same at the beginning and at the end of a reaction.
 - A catalyst usually lowers the activation energy by changing the reaction pathway to avoid the highest-energy state.
 - There is a lower energy barrier for the reaction to occur, so it is more likely that particles will have enough energy to react when they collide.
 - Think of a hiking trail. A catalyst is like a shortcut across the mountain so hikers don’t have to climb all the way to the top.

Quick Fact

For the average reaction occurring at room temperature, raising the temperature by 10 °C approximately doubles the reaction rate.

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



Quick Fact

Car exhaust contains environmental pollutants. Automobiles use catalysts such as palladium and platinum to convert these pollutants into less toxic chemicals and improve air quality.

Changing the activation energy changes the rate of the reaction. However, it does not change the energy of the products or reactants, as seen above, so ΔE remains the same. The rate of a reaction is independent of the overall change in energy.

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 in order to burn fuel, but it does require a lot of energy. There are special catalysts in the body called **enzymes** that help humans burn fuel at normal body temperature.

Without catalysts, the activation energy of the reaction that burns fuel would be too high.

Quick Fact

The slowest known biological reaction would take 1 trillion years without a catalyst. However, this reaction is essential to creating our DNA. With enzymes, the reaction can occur in only 10 milliseconds.

ENERGY OF CHEMICAL REACTIONS

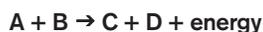
Kinetics deals with the activation energy and the energies along the reaction pathway, but does not describe the difference between the reactants and the products. All chemical reactions involve some overall change in energy, because the products have a different amount of energy than the reactants.

Think of a chemical reaction as a hiking trail over a mountain again. Kinetics deals with how hard it is to climb that mountain and how high the mountain is. The difference in height between the beginning of the trail and the top of the mountain is the activation energy. The energy change of the reaction is the difference in height between the beginning of the trail and the end of the trail. The path a hiker takes to cross the mountain won't change their altitude at the beginning or their altitude at the end. These two altitudes (the beginning and the end) are completely independent of the height of the mountain and the path that a hiker takes to get across.

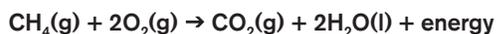
Some chemical reactions require the input of energy and others release energy, depending on whether the products are higher or lower in energy than the reactants.

- **Exothermic reactions:** chemical reactions that produce energy, often in the form of heat, light, or sound.

The reactants themselves lose energy, and that energy is released to the environment around the reaction. The products of an exothermic reaction are more stable (lower in energy) than the reactants.



EXAMPLE:



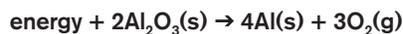
In the exothermic reaction above, methane and oxygen produce carbon dioxide, water, and heat.

- **Endothermic reactions:** chemical reactions that require the input of energy to occur.

The reactants gain energy by taking energy from their surroundings, often in the form of heat or light. An endothermic reaction would feel cold to the touch because it absorbs heat from its environment. The products of an endothermic reaction are less stable (higher in energy) than the reactants.



EXAMPLE:



In the endothermic reaction above, energy is added to bauxite (aluminum oxide) to produce aluminum metal and oxygen gas.

Quick Fact

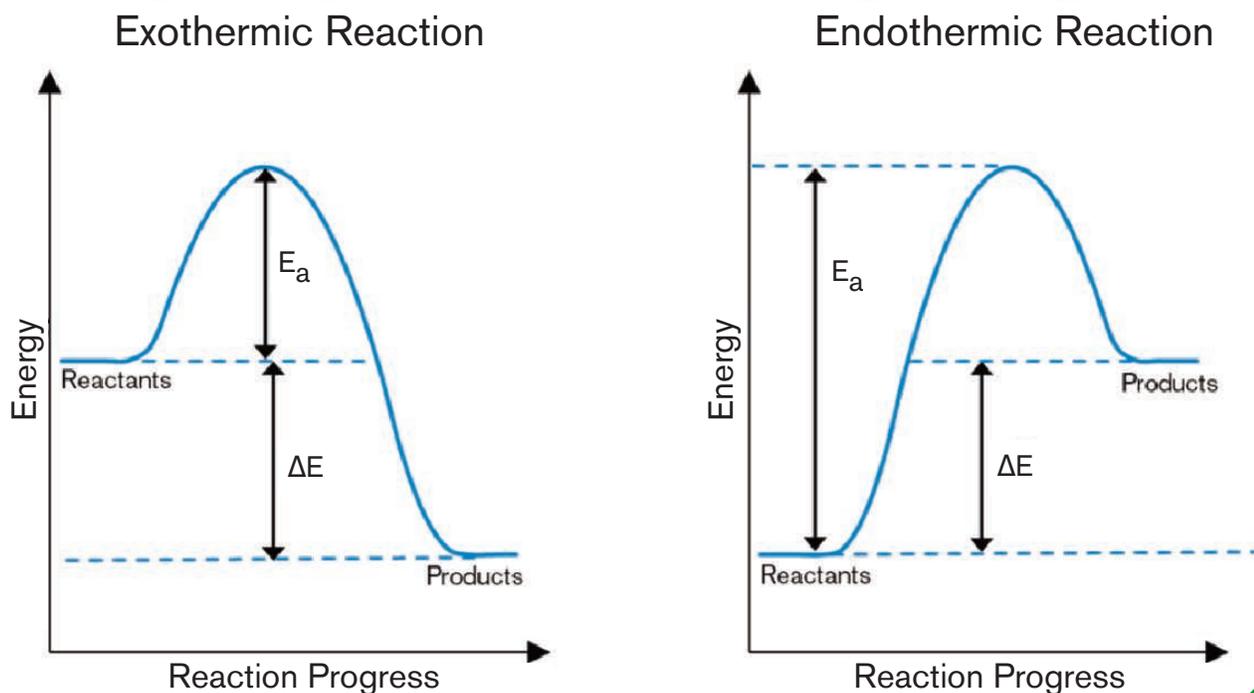
A common example of an exothermic reaction is burning wood in a wood stove. Wood combines with the oxygen in the air to produce carbon dioxide, water, light, and heat.



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.

The reaction coordinate diagrams below show the difference between exothermic and endothermic reactions.



- In the exothermic reaction above, the products are at a lower energy level than the reactants. The difference, ΔE , is the amount of energy released from the reaction.
 - Since $\Delta E = (\text{Energy of the products}) - (\text{Energy of the reactants})$, ΔE is negative. It is lost by the reactants.
- In the endothermic reaction above, the products are at a higher energy level than the reactants. This difference, ΔE , is the amount of energy that has to be put into the reaction.
 - The product energy is greater than the reactant energy, so ΔE is positive. It is gained by the reactants.
- Just because a reaction requires energy in order to start does not mean that it is an endothermic reaction. The exothermic reaction above requires a certain amount of energy (E_a) to get over the energy barrier, but releases more energy during the reaction than it takes to get started. Chemists often initiate exothermic reactions, adding a small amount of energy, to produce much more energy.

Quick Fact

Diamond and graphite are both forms of pure carbon. Graphite is more stable than diamond at normal room temperature and pressure, so the reaction of diamond turning into graphite is exothermic and should spontaneously occur. The reason that diamonds do not turn into graphite is because the activation energy for the reaction is extremely high, making it almost impossibly difficult for the reaction to occur.

Think About It...

What does the reaction coordinate diagram for diamond turning into graphite look like? How do the energy of the reactants, the energy of the products, and activation energy compare to one another?

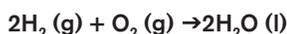
GENERAL TYPES OF CHEMICAL REACTIONS

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

SYNTHESIS (OR COMBINATION) REACTION: a chemical reaction in which smaller molecules combine to form a larger one. As shown below, two or more reactants (A and B) combine to form a product (AB).



EXAMPLE:

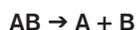


The (g) and (l) after the compounds correspond to the phases of each of those compounds. A (g) means that compound is a gas, (l) means a liquid, and an (s) means a solid. You might also see (aq), which stands for aqueous solution – or a liquid mixture of a compound where water is the solvent.

Think About It...

The reaction to the left could be used for hydrogen-powered cars. There is plenty of oxygen in the air to fuel the reaction, but where would the hydrogen come from?

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



EXAMPLE:

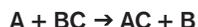


In this decomposition reaction, mercury(II) oxide (HgO) splits into liquid mercury and oxygen gas. The small triangle above the arrow means the reaction needs energy (usually in the form of heat) to take place.

Quick Fact

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

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



EXAMPLE:

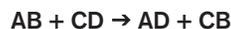


This displacement reaction happens when a piece of solid (s) magnesium metal is combined with some aqueous (aq) hydrochloric acid. When these reactants combine, they produce two products: a liquid solution called aqueous magnesium chloride and hydrogen gas (g).

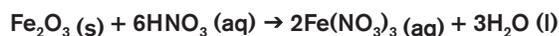
Think About It...

How would you describe the displacement reaction to the left?

DOUBLE DISPLACEMENT (OR REPLACEMENT) REACTION: a chemical reaction between two compounds, AB and CD, in which part of one reactant trades places with part of another reactant. The elements are rearranged to form two or more different compound products, AD and CB.



EXAMPLE:



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

Quick Fact

The reaction shown 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 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 at least one atom's oxidation number changes. This change is the result of an atom or group of atoms gaining or losing electrons. All oxidation-reduction reactions involve the transfer of electrons between substances.

- **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:



The reaction between Na and Cl₂ gas is a redox reaction. The complete reaction can be divided in half to illustrate the oxidation and reduction parts.

Oxidation: The sodium metal atoms begin with an oxidation number of zero and end with an oxidation number of +1. Each sodium metal atom loses one electron, and is oxidized to a sodium cation, Na⁺.



Reduction: Each chlorine atom in Cl₂ gas begins with an oxidation number of zero and ends with an oxidation number of -1. Each chlorine atom gains one electron, and is reduced to a chloride anion, Cl⁻.



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 atoms oxidize the sodium metal atoms to sodium cations, so chlorine is the oxidizing agent. On the other hand, sodium metal atoms reduce the chlorine atoms to chloride anions, so sodium is called the reducing agent. During a redox reaction, the oxidizing agent is reduced and the reducing agent is oxidized.

Quick Fact

An easy way to remember the different redox reactions is the acronym "OIL RIG":
Oxidation Is electron **L**oss,
Reduction Is electron **G**ain.

Redox reactions occur all around us. The combination of hydrogen and chlorine gas to form hydrogen chloride 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 with oxygen as a reactant 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.

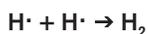
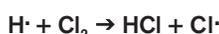
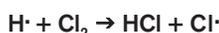


CHAIN REACTIONS

A **chain reaction** is a series of chemical reactions in which the products of one reaction initiate further 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. Chain reactions are found in gas explosions, combustion, the formation of smog, and nuclear reactions.

- **Initiation reaction:** the chemical reaction that starts a chain reaction. The product sets up a sequence of repeating reactions.
- **Propagating reactions:** reactions whose products cause other reactions.
- **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) and the reaction stops.

EXAMPLE:



Quick Fact

Most chain reactions occur very quickly. Sometimes, the rate of the reaction increases so much that it results in an explosion.

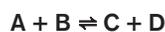
When chlorine and hydrogen interact in the presence of light, a chain reaction occurs.

- **The light absorbed by a Cl_2 molecule breaks the molecule into two separate atoms, called chlorine free radicals ($\text{Cl}\cdot$). “Free radical” means that it has an unpaired valence electron.**
- **The chlorine radicals are very reactive, and react rapidly with hydrogen molecules to produce hydrogen chloride and hydrogen free radicals ($\text{H}\cdot$).**
- **The hydrogen radicals react with Cl_2 molecules, producing hydrogen chloride and more 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, a termination reaction occurs when two chlorine radicals combine with each other to form Cl_2 , which is a stable molecule known as chlorine gas.**

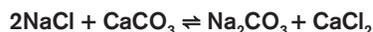
REVERSIBLE REACTIONS AND EQUILIBRIUM

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. Under certain conditions, a reversible reaction may go fully from reactants to products, using up all of the reactants. Under other conditions, the reaction may not go fully to completion and may not use up all of the reactants. If a reaction goes to completion, that means it has used up all of one or more of the reactants, so no more product can be made. Not all reactions go to completion.

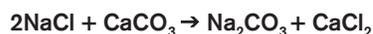
- Reversible reactions are usually represented in a chemical equation by a double arrow:



EXAMPLE:

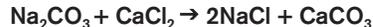


The forward reaction can be seen in saltwater lakes:



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

The reverse reaction is as follows:



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 naturally occur.

- In a reversible reaction, both reactants and products may be present at the same time in a state of equilibrium.

Equilibrium: the state of a chemical reaction in which the forward and reverse reactions occur at the same rate. As a result, the concentrations of the reactants and the products do not change (unless the conditions change). However, the concentrations of the reactants and products do not have to be equal to each other.

Equilibrium also describes how far a reversible reaction will go towards completion under given conditions. It indicates how much product a reaction will actually produce, rather than the maximum amount of product that could be made from the amount of reactants. Scientists can change the equilibrium, and the amount of product formed, by manipulating conditions of the reaction such as temperature and concentration.

Quick Fact

French chemist Claude Berthollet came up with the idea of reversible reactions while visiting the Natron Lakes in Egypt as part of Napoléon Bonaparte's military expedition. He noticed that both the products (Na_2CO_3) and the reactants (NaCl and CaCO_3) of the reaction above were present at a saltwater lake with limestone rock, and went on to develop the idea of chemical equilibrium.

SECTION IV: BALANCING CHEMICAL EQUATIONS

OBJECTIVES

- Describe the law of conservation of mass.
- Apply the law of conservation of mass 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 a reaction. A chemical equation uses chemical symbols to show the reactants and products involved. A chemical equation can also show the physical states of the reactants and products, as well as the involvement of heat, light, or radiation.

Unlike mathematical equations, the two sides are separated by an arrow to show that the reactants form the products.

CONSERVATION OF MASS

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.)

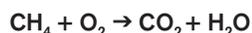
Because all matter is made of atoms, the law implies the conservation of atoms as well. Therefore, although atoms may be rearranged, they are not lost.

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. It also shows the lowest whole-number ratio of products to reactants.

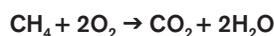
- If a hydrogen atom goes into a reaction, it has to appear somewhere in the products of the reaction.
- Likewise, if three hydrogen atoms appear on the reactant side of a chemical equation, three must appear on the product side.

EXAMPLE:

Methane interacts with oxygen in air as shown in the reaction below:



As written above, the equation correctly indicates that methane and oxygen combine to form carbon dioxide and water. However, this reaction violates conservation of matter. There are more oxygen atoms on the right than on the left, and more hydrogen atoms on the left than the right. The correctly balanced equation for the reaction looks like this:



Quick Fact

Diatomic elements must always have a subscript of 2 in pure form. In the reaction shown to the left, oxygen must be present as O_2 and not just O . The diatomic elements are Br, I, N, Cl, H, O, and F.

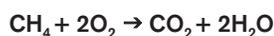
To determine the number of atoms of one element in a chemical formula:

- Multiply the number in front of the chemical formula by the subscript number on that atom in the chemical formula. The number in front of the chemical formula is called a **coefficient**. It indicates how many of that molecule or atom there are. The coefficient of a molecule applies to all of the elements in the molecule's formula.
- The number one is never written. CH_4 stands for $1\text{C}_1\text{H}_4$.
- 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:



Reactant Side of Equation

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

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

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

Product Side of Equation

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

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

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

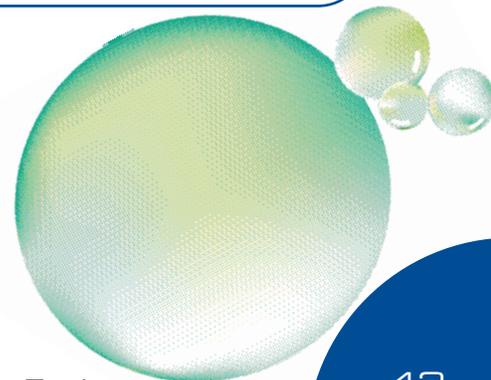
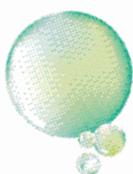
The number of each type of atom is the same in the reactants as it is in the products, therefore the reaction is balanced.

HISTORY: ANTOINE LAVOISIER (1743-1794)

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

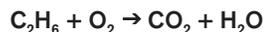
Lavoisier is known as the father of modern chemistry. He changed chemistry from a qualitative to a quantitative science.

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



BALANCING CHEMICAL EQUATIONS

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



Reactant Side of Equation

$$\text{C: } 1 \times 2 = 2$$

$$\text{H: } 1 \times 6 = 6$$

$$\text{O: } 1 \times 2 = 2$$

Product Side of Equation

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

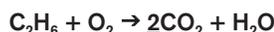
$$\text{H: } 1 \times 2 = 2$$

$$\text{O: } (1 \times 2) + (1 \times 1) = 3$$

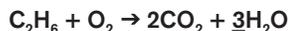
- There is 1 fewer carbon atom, 4 fewer hydrogen atoms, and 1 more oxygen atom in the products than there are in the reactants. This equation is not balanced.

2. Balance the equation. An easy way to start is to balance the elements that appear in only one reactant and one product (carbon and hydrogen in this reaction). Once those elements are balanced, move on to the elements that appear in multiple reactants or products (oxygen in this reaction). To test each different equation multiply the different atoms and molecules on each side by different amounts based on their coefficients.

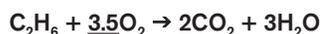
- To get the same number of carbon atoms on both sides, multiply CO_2 (on the right side) by 2. This is shown by placing a 2 in front of CO_2 .



- 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.** Changing the subscripts changes the ratio of different elements in the compound, which makes it a different molecule altogether. In this case, 1 molecule of CO_2 is made of 1 carbon atom and 2 oxygen atoms. The coefficient says that the reaction produces 2 molecules of CO_2 .
- To get the same number of hydrogen atoms on both sides, add a coefficient of 3 in front of H_2O on the right side, making it $3\text{H}_2\text{O}$.



- Now there are 2 carbon atoms and 6 hydrogen atoms on each side. However, the number of oxygen atoms is not the same. There are 2 on the left side and 7 on the right side (4 from CO_2 , 3 from H_2O).
- Add a coefficient of 3.5 in front of O_2 on the left side, making it 3.5O_2 .



This equation has the same number of each type of atom on both sides, but is not quite done.

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



This equation is the properly balanced equation for the reaction.

Quick Fact

Often the best way to balance an equation is by trial and error, and testing a few different numbers until one set of coefficients works.

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.

ACIDS

Substances that release hydrogen ions (H^+) when dissolved in water are **acids**. An acidic solution has an excess of hydrogen ions (H^+).

- Acids often donate H^+ ions to bases.
- Acids dissociate (break down) into ions in solution, making them able to conduct electricity. Some acids are corrosive in nature, and have the ability to dissolve some metals.
 - Most acids can react with a metal to produce 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.



Quick Fact

The word “acid” comes from the Latin term “acidus,” which means sour. Substances with a sour taste (like lemon juice) are usually acidic.



BASES

Substances that either release hydroxide ions (OH^-) when dissolved in water or that accept H^+ ions from acids are **bases**. A basic solution has an excess of OH^- ions.

- Bases often 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 harm human skin.

- Although the term “alkali” is often used as a synonym for base, they are not the same thing. **Alkalis** are a type of base made of ionic salts of an alkali metal or an alkaline earth metal. *All alkalis are bases, but not all bases are alkalis.*

EXAMPLE:

Calcium carbonate (CaCO₃) and soda lye (NaOH) are bases that are also alkali salts. Ammonia (NH₃) is a base but *not* an alkali.

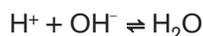
Quick Fact

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

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 lot of hydrogen ions are released into a solution (high concentration of hydrogen ions), that solution 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. pH values do not have units.
 - Substances with a pH value less than 7 are acids. The smaller the number on the pH scale, the more acidic a substance is. A substance with a pH of 1 is a stronger acid than a substance with a pH of 5 (though both are acidic).
 - Substances with a pH value greater than 7 are bases. The higher the number on the pH scale, the more basic a substance is. A substance with a pH of 13 is a stronger base than a substance with a pH of 10 (though both are basic).
 - 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.



- Small changes on the pH scale actually mean large changes in acidity. An increase of just one pH unit indicates that the concentration of H⁺ ions (and the acidity) has increased by a factor of ten. For example, the concentration of H⁺ ions in a solution with a pH of 5.0 is ten times that in a solution with a pH of 6.0; the solution with pH 5.0 is ten times more acidic. Similarly, a solution with a pH of 3.0 is 1,000 times more acidic than a solution with a pH of 6.0.

Quick Fact

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 less than zero.

Quick Fact

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

The table below lists some common acids and bases on the pH scale. All of the substances whose pH is less than the pH of water (pictured above water on the table) are acidic. All of the substances whose pH is greater than the pH of water are basic.

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

*See the subsection on [Indicators](#).

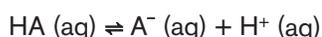
STRENGTH OF ACIDS AND BASES

Acids and bases may be strong or weak depending on how well an acid or base produces ions in water. Many things will affect the strength of an acid or a base.

ACIDS

- 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, HA, dissociating (producing hydrogen ions) is shown below:



- For some acids, the strength of an acid may be affected by the size of the anion produced when the hydrogen is released into water (see the section on Ions from *The Core of Chemistry*). **Within a group, acid strength increases as atomic radius increases moving down the periodic table.**

Think About It...

Like acids, metallic substances can also conduct electricity. What do they have in common? How do acids conduct electricity differently than metals?

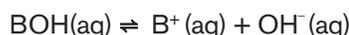


– Larger anions are more stable, so they are more easily separated from the hydrogen ion and the corresponding acid is stronger (produces more H^+ ions). Hydroiodic acid ($\text{HI}_{(\text{aq})}$) is a stronger acid than hydrofluoric acid ($\text{HF}_{(\text{aq})}$) because the iodide ion ($\text{I}^-_{(\text{aq})}$) is larger than the fluoride ion ($\text{F}^-_{(\text{aq})}$), and therefore more stable. Binary acids (made from hydrogen and one other element) increase in strength going down a group on the periodic table.

- Electronegativity also affects acid strength (see the subsection on **Electronegativity**). A more electronegative atom more strongly attracts electrons away from hydrogen atoms in solution, so more hydrogen atoms lose an electron to become H^+ cations. The more electronegative the element bonded to hydrogen in an acid, the stronger the acid. For example, hydrofluoric acid ($\text{HF}_{(\text{aq})}$) is more acidic than ammonia ($\text{NH}_{3(\text{aq})}$) because fluorine is more electronegative than nitrogen. **Within a period, acid strength increases as electronegativity increases moving from left to right across the periodic table.**
- Strong acids include hydroiodic acid (HI), hydrobromic acid (HBr), hydrochloric acid (HCl), sulfuric acid (H_2SO_4), and nitric acid (HNO_3).

BASES

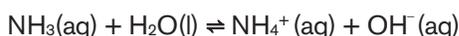
- Strong bases act in a manner similar to strong acids, producing hydroxide ions instead of hydrogen ions. The equation of a base, BOH, dissociating looks like this:



- Strong bases include sodium hydroxide (NaOH), potassium hydroxide (KOH), and lithium hydroxide (LiOH).
- Bases can also produce OH^- ions by accepting H^+ ions. In the equation below, a base, B, accepts an H^+ ion from water:



- Ammonia, NH_3 , is an example of a base that accepts H^+ ions, as shown in the equation below:

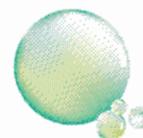


INDICATORS

Indicators are substances that change color at a specific pH. They provide a way to determine approximately how acidic or basic a solution is. Some common indicators:

- **Litmus paper** turns red in an acidic solution and blue in a basic solution.
- **Phenolphthalein solution** changes from colorless to pink for a pH greater than 8.2.
- **Bromothymol Blue (BTB)** turns yellow in acidic solutions and blue in basic solutions.

To get more accurate, quantitative pH measurements, scientists use an instrument called a pH meter.



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.

SECTION VI: RADIOACTIVITY AND 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 human-made elements and their location on the periodic table.



Elements tend to exist in multiple forms, called isotopes. Remember, isotopes are atoms of the same element that contain different numbers of neutrons. Isotopes of the same element have the same atomic number, but different mass numbers (see the section on Isotopes from *The Core of Chemistry*).

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

EXAMPLE:

Hydrogen can exist as one of three isotopes (protium, deuterium, and tritium). The lightest isotope, protium, tends to undergo chemical reactions at the fastest rate.

RADIOACTIVITY

Radioactivity is the spontaneous breakdown of an unstable nucleus in an atom. A radioactive atom decays to produce a lighter atom (with fewer neutrons and/or protons) and releases 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.

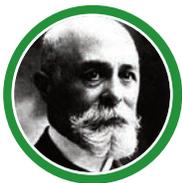
Quick Fact

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. It takes 5,730 years for half of the nuclei in a sample of a carbon-14 to 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.

HISTORY: HENRI BECQUEREL (1852-1908)



In 1896, French physicist Henri Becquerel accidentally discovered radioactivity while investigating phosphorescence in uranium salts.

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 Schindt 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.



POLONIUM

Atomic #84

Polonium was discovered by Marie and Pierre Curie in 1898. 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 lead-206 and a lot of energy (140 watts per gram).



RADON

Atomic #86

Radon is produced by 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 soil and rocks and can sometimes build up in homes.

The World Health Organization estimates that radon is responsible for up to 14% of all lung cancer cases. Radon test kits are available to check for radon accumulation in homes, especially basement levels.



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.

ALPHA (α) RADIATION (ALPHA PARTICLES): radiation composed of helium-4 nuclei. Alpha particles have a nucleus that is the same as helium: two protons and two neutrons. When an atom undergoes alpha decay, it loses two protons and two neutrons to form an alpha particle. The atom's atomic number decreases by two, and its mass number decreases by four.

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

EXAMPLE:

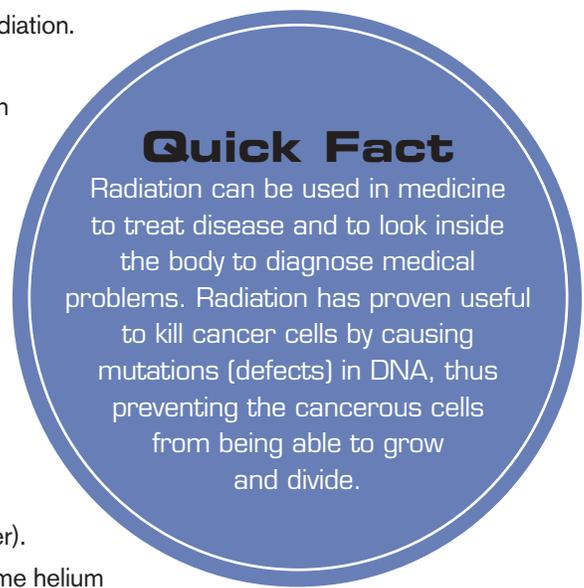
Uranium-238 has 92 protons and 146 neutrons. When uranium-238 undergoes alpha decay, it loses two protons and two neutrons to produce an alpha particle and thorium-234 (90 protons and 144 neutrons).

BETA (β) RADIATION (BETA PARTICLES): radiation composed of high-velocity electrons emitted from an unstable nucleus.

- 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.
- Sometimes the release of a beta particle is not enough to get rid of the extra energy in an unstable nucleus. In this case, the nucleus often releases the rest of the excess energy in the form of gamma rays.

GAMMA (γ) RADIATION (GAMMA RAYS): 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.



Quick Fact

Radiation can 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.

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, vast amounts of energy are released.

- Uranium nuclei can be split easily by bombarding them with neutrons.
- Once a uranium nucleus is split, multiple neutrons are released. Each of these neutrons initiates other fission reactions, resulting in a chain reaction.

Quick Fact

The atomic bomb developed by the United States during World War II used a nuclear fission reaction beginning with the radioactive isotope uranium-235. Nuclear fission is also used in nuclear power plants to generate energy.



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 glass interesting yellow and green colors and a fluorescent effect.

NUCLEAR FUSION: a nuclear reaction that occurs when the nuclei of two 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 energy (in the form of heat and light) through nuclear fusion. In the sun, hydrogen nuclei fuse to make helium.

Quick Fact

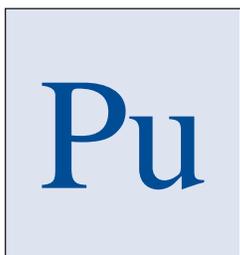
The hydrogen bomb uses nuclear fusion. Hydrogen nuclei fuse to form helium. In the process, they release huge amounts of energy and create a massive explosion.

HISTORY: ENRICO FERMI (1901-1954)



Enrico Fermi was an Italian physicist most known for his work on beta decay, the development of the first nuclear reactor, and 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 (fermium) was named after him.



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 the United States 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.

HUMAN-MADE ELEMENTS

All of the elements with atomic numbers greater than 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 human-made, transuranic elements are named for important chemists or physicists. Curium (atomic number 96), for example, is named after Marie Curie and her husband, Pierre Curie. Other human-made elements are listed in the table below:

| 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 modern 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 is known for her work on the discovery of nuclear fission. |

Quick Fact

In December 2015, the International Union of Pure and Applied Chemistry (IUPAC) verified the discovery of four new transuranic elements. The addition of elements 113, 115, 117, and 118 completes the seventh row of the periodic table.

Quick Fact

The new elements were named nihonium (element 113), moscovium (115), tennessine (117), and oganesson (118). Nihonium is the first element to be named by researchers in Asia—the common Japanese name for Japan is "nihon."



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