## Overview of Atomic Structure

Atoms are made up of particles called protons, neutrons, and electrons, which are responsible for the mass and charge of atoms.

#### Key Points

* An atom is composed of two regions: the nucleus, which is in the center of the atom and contains protons and neutrons, and the outer region of the atom, which holds its electrons in orbit around the nucleus.
* Protons and neutrons have approximately the same mass, about 1.67 × 10-24 grams, which scientists define as one atomic mass unit (amu) or one Dalton.
* Each electron has a negative charge (-1) equal to the positive charge of a proton (+1).
* Neutrons are uncharged particles found within the nucleus.

#### Key Terms

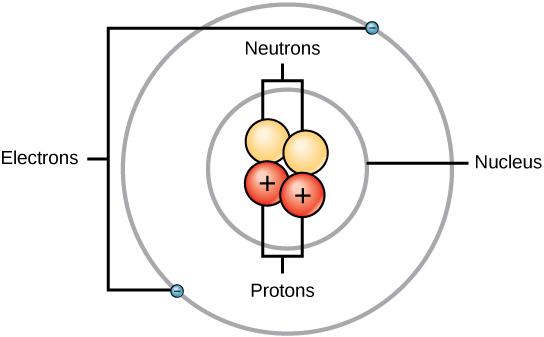
* **atom**: The smallest possible amount of matter which still retains its identity as a chemical element, consisting of a nucleus surrounded by electrons.
* **proton**: Positively charged subatomic particle forming part of the nucleus of an atom and determining the atomic number of an element. It weighs 1 amu.
* **neutron**: A subatomic particle forming part of the nucleus of an atom. It has no charge. It is equal in mass to a proton or it weighs 1 amu.

An atom is the smallest unit of matter that retains all of the chemical properties of an element. Atoms combine to form molecules, which then interact to form solids, gases, or liquids. For example, water is composed of hydrogen and oxygen atoms that have combined to form water molecules. Many biological processes are devoted to breaking down molecules into their component atoms so they can be reassembled into a more useful molecule.

### Atomic Particles

Atoms consist of three basic particles: protons, electrons, and neutrons. The nucleus (center) of the atom contains the protons (positively charged) and the neutrons (no charge). The outermost regions of the atom are called electron shells and contain the electrons (negatively charged). Atoms have different properties based on the arrangement and number of their basic particles.

The hydrogen atom (H) contains only one proton, one electron, and no neutrons. This can be determined using the atomic number and the mass number of the element (see the concept on atomic numbers and mass numbers).



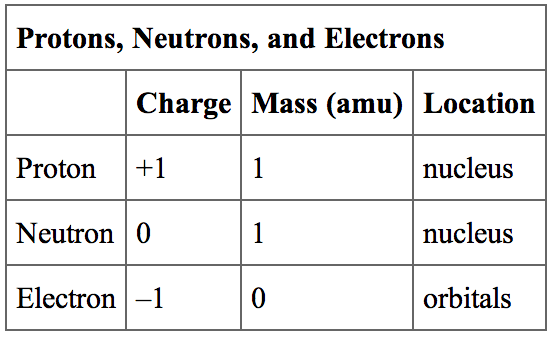
**Structure of an atom**: Elements, such as helium, depicted here, are made up of atoms. Atoms are made up of protons and neutrons located within the nucleus, with electrons in orbitals surrounding the nucleus.

### Atomic Mass

Protons and neutrons have approximately the same mass, about 1.67 × 10-24 grams. Scientists define this amount of mass as one atomic mass unit (amu) or one Dalton. Although similar in mass, protons are positively charged, while neutrons have no charge. Therefore, the number of neutrons in an atom contributes significantly to its mass, but not to its charge.

Electrons are much smaller in mass than protons, weighing only 9.11 × 10-28 grams, or about 1/1800 of an atomic mass unit. Therefore, they do not contribute much to an element’s overall atomic mass. When considering atomic mass, it is customary to ignore the mass of any electrons and calculate the atom’s mass based on the number of protons and neutrons alone.

Electrons contribute greatly to the atom’s charge, as each electron has a negative charge equal to the positive charge of a proton. Scientists define these charges as “+1” and “-1. ” In an uncharged, neutral atom, the number of electrons orbiting the nucleus is equal to the number of protons inside the nucleus. In these atoms, the positive and negative charges cancel each other out, leading to an atom with no net charge.



**Protons, neutrons, and electrons**: Both protons and neutrons have a mass of 1 amu and are found in the nucleus. However, protons have a charge of +1, and neutrons are uncharged. Electrons have a mass of approximately 0 amu, orbit the nucleus, and have a charge of -1.

**Exploring Electron Properties**: Compare the behavior of electrons to that of other charged particles to discover properties of electrons such as charge and mass.

### Volume of Atoms

Accounting for the sizes of protons, neutrons, and electrons, most of the volume of an atom—greater than 99 percent—is, in fact, empty space. Despite all this empty space, solid objects do not just pass through one another. The electrons that surround all atoms are negatively charged and cause atoms to repel one another, preventing atoms from occupying the same space. These intermolecular forces prevent you from falling through an object like your chair.

## Atomic Number and Mass Number

The atomic number is the number of protons in an element, while the mass number is the number of protons plus the number of neutrons.

#### Key Points

* Neutral atoms of each element contain an equal number of protons and electrons.
* The number of protons determines an element’s atomic number and is used to distinguish one element from another.
* The number of neutrons is variable, resulting in isotopes, which are different forms of the same atom that vary only in the number of neutrons they possess.
* Together, the number of protons and the number of neutrons determine an element’s mass number.
* Since an element’s isotopes have slightly different mass numbers, the atomic mass is calculated by obtaining the mean of the mass numbers for its isotopes.

#### Key Terms

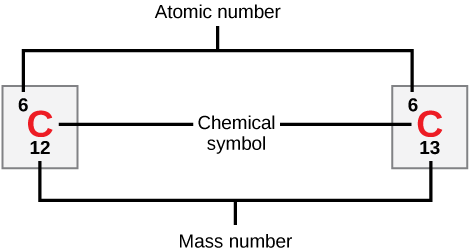
* **mass number**: The sum of the number of protons and the number of neutrons in an atom.
* **atomic number**: The number of protons in an atom.
* **atomic mass**: The average mass of an atom, taking into account all its naturally occurring isotopes.

### Atomic Number

Neutral atoms of an element contain an equal number of protons and electrons. The number of protons determines an element’s atomic number (Z) and distinguishes one element from another. For example, carbon’s atomic number (Z) is 6 because it has 6 protons. The number of neutrons can vary to produce isotopes, which are atoms of the same element that have different numbers of neutrons. The number of electrons can also be different in atoms of the same element, thus producing ions (charged atoms). For instance, iron, Fe, can exist in its neutral state, or in the +2 and +3  ionic states.

### Mass Number

An element’s mass number (A) is the sum of the number of protons and the number of neutrons. The small contribution of mass from electrons is disregarded in calculating the mass number. This approximation of mass can be used to easily calculate how many neutrons an element has by simply subtracting the number of protons from the mass number. Protons and neutrons both weigh about one atomic mass unit or amu. Isotopes of the same element will have the same atomic number but different mass numbers.



**Atomic number, chemical symbol, and mass number**: Carbon has an atomic number of six, and two stable isotopes with mass numbers of twelve and thirteen, respectively. Its average atomic mass is 12.11.

Scientists determine the atomic mass by calculating the mean of the mass numbers for its naturally-occurring isotopes. Often, the resulting number contains a decimal. For example, the atomic mass of chlorine (Cl) is 35.45 amu because chlorine is composed of several isotopes, some (the majority) with an atomic mass of 35 amu (17 protons and 18 neutrons) and some with an atomic mass of 37 amu (17 protons and 20 neutrons).

Given an atomic number (Z) and mass number (A), you can find the number of protons, neutrons, and electrons in a neutral atom. For example, a lithium atom (Z=3, A=7 amu) contains three protons (found from Z), three electrons (as the number of protons is equal to the number of electrons in an atom), and four neutrons (7 – 3 = 4).

## Isotopes

Isotopes are various forms of an element that have the same number of protons, but a different number of neutrons.

#### Key Points

* Isotopes are atoms of the same element that contain an identical number of protons, but a different number of neutrons.
* Despite having different numbers of neutrons, isotopes of the same element have very similar physical properties.
* Some isotopes are unstable and will undergo radioactive decay to become other elements.
* The predictable half-life of different decaying isotopes allows scientists to date material based on its isotopic composition, such as with Carbon-14 dating.

#### Key Terms

* **isotope**: Any of two or more forms of an element where the atoms have the same number of protons, but a different number of neutrons within their nuclei.
* **half-life**: The time it takes for half of the original concentration of an isotope to decay back to its more stable form.
* **radioactive isotopes**: an atom with an unstable nucleus, characterized by excess energy available that undergoes radioactive decay and creates most commonly gamma rays, alpha or beta particles.
* **radiocarbon dating**: Determining the age of an object by comparing the ratio of the 14C concentration found in it to the amount of 14C in the atmosphere.

### What is an Isotope?

Isotopes are various forms of an element that have the same number of protons but a different number of neutrons. Some elements, such as carbon, potassium, and uranium, have multiple naturally-occurring isotopes. Isotopes are defined first by their element and then by the sum of the protons and neutrons present.

* Carbon-12 (or 12C) contains six protons, six neutrons, and six electrons; therefore, it has a mass number of 12 amu (six protons and six neutrons).
* Carbon-14 (or 14C) contains six protons, eight neutrons, and six electrons; its atomic mass is 14 amu (six protons and eight neutrons).

While the mass of individual isotopes is different, their physical and chemical properties remain mostly unchanged.

Isotopes do differ in their stability. Carbon-12 (12C) is the most abundant of the carbon isotopes, accounting for 98.89% of carbon on Earth. Carbon-14 (14C) is unstable and only occurs in trace amounts. Unstable isotopes most commonly emit alpha particles (He2+) and electrons. Neutrons, protons, and positrons can also be emitted and electrons can be captured to attain a more stable atomic configuration (lower level of potential energy ) through a process called radioactive decay. The new atoms created may be in a high energy state and emit gamma rays which lowers the energy but alone does not change the atom into another isotope. These atoms are called radioactive isotopes or radioisotopes.

**Calculating Average Atomic Mass**

**Key Points**

* An element can have differing numbers of neutrons in its nucleus, but it always has the same number of protons. The versions of an element with different neutrons have different masses and are called isotopes.
* The average atomic mass for an element is calculated by summing the masses of the element’s isotopes, each multiplied by its natural abundance on Earth.
* When doing any mass calculations involving elements or compounds, always use average atomic mass, which can be found on the periodic table.

**Key Terms**

* **mass number**: The total number of protons and neutrons in an atomic nucleus.
* **natural abundance**: The abundance of a particular isotope naturally found on the planet. Described in units of %.
* **average atomic mass**: The mass calculated by summing the masses of an element’s isotopes, each multiplied by its natural abundance on Earth.

The atomic number of an element defines the element’s identity and signifies the number of protons in the nucleus of one atom. For example, the element hydrogen will always have one proton in its nucleus. The element helium will always have two protons in its nucleus.

**Isotopes**

Atoms of the same element can, however, have differing numbers of neutrons in their nucleus. For example, stable helium atoms exist that contain either one or two neutrons, but both atoms have two protons. These different types of helium atoms have different masses (3 or 4 atomic mass units), and they are called isotopes.

For any given isotope, the sum of the numbers of protons and neutrons in the nucleus is called the **mass number.** This is because each proton and each neutron weigh one atomic mass unit (amu) relatively speaking. By adding together, the number of protons and neutrons and multiplying by 1 amu, you can calculate the mass of the atom. However, the actual atomic mass of any isotope is not a perfect whole number because the mass of every atom is determined relative to the mass of Carbon-12. Only Carbon-12 has a mass of exactly 12.000 amu. Therefore you will find the masses of individual isotopes to be slightly off from whole numbers, but only by a small degree.

All elements exist as a collection of isotopes. The word ‘isotope’ comes from the Greek ‘isos’ (meaning ‘same’) and ‘topes’ (meaning ‘place’) because the elements can occupy the same place on the periodic table while being different in subatomic construction.

**Calculating Average Atomic Mass**

The average atomic mass of an element is the sum of the masses of its isotopes, each multiplied by its natural abundance (the decimal associated with percent of atoms of that element that are of a given isotope).

The average atomic mass of an element can be found on the periodic table, typically under the elemental symbol. When data are available regarding the natural abundance of various isotopes of an element, it is simple to calculate the average atomic mass.

* For helium, there is approximately one isotope of Helium-3 for every million isotopes of Helium-4; therefore, the average atomic mass is very close to 4 amu (4.002602 amu).
* Chlorine consists of two major isotopes, one with 18 neutrons (75.76 percent of natural chlorine atoms) and one with 20 neutrons (24.24 percent of natural chlorine atoms). The atomic number of chlorine is 17 (it has 17 protons in its nucleus).

The average atomic mass of an element is the sum of the masses of its isotopes, each multiplied by its natural abundance. This means you multiple the % of each element by its atomic mass.

To calculate the average mass, first convert the percentages into decimals (divide them by 100). Then, calculate the mass numbers. The chart below shows the exact atomic masses and natural abundances of the two naturally occurring chloring isotopes.

|  |  |  |
| --- | --- | --- |
| **Isotopes of Chlorine** | **Atomic mass (amu)** | **Natural Abundance** |
| Chlorine-35 | 34.96885 | 75.76% |
| Chlorine-37 | 36.965902 | 24.24% |

You multiply 34.96885 x 0.7576 = 26.49 amu

You multiply 36.965902 x 0.2424 = 8.961 amu

Average Atomic Mass of Chlorine is 35.45 amu

See if you can calculate Boron’s average atomic mass from the data below.

|  |  |  |
| --- | --- | --- |
| **Isotopes of Chlorine** | **Atomic mass (amu)** | **Natural Abundance** |
| Boron-10 | 10.012937 | 19.99% |
| Boron-11 | 11.009305 | 80.01% |

**Parts of the Periodic Table**

**Categories Location**

**Metals**  found to the left of the staircase on the periodic table

**Nonmetals** found to the right of the staircase on the periodic table

**Metalloids** touching two lines of the staircase on the periodic table

**Noble Gases** found in Group 18 of the periodic table

**Exceptions:**  Hydrogen (H) is on the left side of the table but is a nonmetal

Aluminum (Al) and Polonium (Po) are metals even though they touch two lines on the staircase.

**Categories Physical Properties**

**Metals**  luster (shiny), malleable, ductile, high melting points, solids at room temperature, conductive

**Nonmetals** dull, brittle, low melting points/boiling points, non-conductive, found in all three phases at room temperature

**Metalloids** contains a mix of both metal and non-metal properties

**Noble Gases** found as gases at room temperature

**Exceptions:**  Mercury (Hg) is the only liquid metal at room temperature

Bromine (Br) is the only liquid nonmetal at room temperature, making Group 17 (Halogens) the only group to contain all three phases of matter.

**Categories Chemical Properties**

**Metals**  react with nonmetals; metals get more reactive as you go down and to the left of the periodic table. **Francium (Fr)** is the most most reactive metal. It is the most **metallic.**

**Nonmetals** react with both metals and nonmetals; nonmetals get more reactive as you go up and to the right of the periodic table. **Fluorine (F)** is the most reactive nonmetal. It is the most **non-metallic.**

**Metalloids** react with either metals or nonmetals but it depends on the metalloid.

**Noble Gases** Unreactive

**Exceptions:**  Fluorine is so reactive that it can even react with certain Noble Gases.

**Groups Group Names Special Properties**

**1 Alkali metals** Never found alone in nature (too reactive), can be made from electrolysis of a salt or other chemical reaction. React violently with water, most reactive metals

**2 Alkaline Earth** Never found alone in nature (too reactive), Second most reactive metals

**3 - 11 Transition metals** Can be found alone in nature; salts made with transition metals have colors

**17 Halogens** Never found alone in nature, all are diatomic, most reactive nonmetals

**18 Noble gases** Always found alone in nature, Unreactive

**Salts** are compounds *usually* made from metals and nonmetals. (There are some exceptions).

**Allotropes** are versions of the *same element* but with a *different molecular structure*. This gives them both different physical and chemical properties. Examples: O2 (oxygen) and O3 (ozone)

Carbon can be found as, charcoal, graphite, diamond and Buckminster fullerene

**Naming Ionic Compounds**

An ionic compound is named first by its cation and then by its anion.

**Key Points**

* Most cations and anions can combine to form neutral compounds (typically solids under normal conditions) that are usually referred to as **salts.**
* The net charge of an ionic compound must be **zero**. Therefore, the number of cations and anions in an ionic compound must be balanced to make an electrically neutral molecule.
* When naming ionic compounds, the cation retains the same name as the element. The anion’s name is similar to the elemental name, but the ending of the name has been removed and replaced with **“-ide.”**
* If a metallic element has cations of different charges, which cation is used has to be indicated by its suffix (an older method) or by **Roman numerals** in parentheses after its name in writing (**the Stock system** ).

**Key Terms**

* **Stock system**: A system of naming that includes using Roman numerals to indicate the charge on transition metals.

In chemistry, an ionic compound is a chemical compound in which ions are held together by ionic bonds. Usually, the positively charged portion consists of metal cations and the negatively charged portion is an anion or polyatomic ion. Ionic compounds have high melting and boiling points, and they tend to be hard and brittle.

Ions can be single atoms, as the sodium and chlorine in common table salt (sodium chloride), or more complex (**polyatomic**) groups such as the carbonate in calcium carbonate. But to be considered an ion, they must carry a positive or negative charge. Thus, in an ionic bond, one ‘bonder’ must have a positive charge and the other a negative one. By sticking to each other, they resolve, or partially resolve, their separate charge imbalances. Positive to positive and negative to negative ionic bonds do not occur.

Most cations and anions can combine to form solid compounds that are usually known as salts. The one overriding requirement is that the resulting compound must be electrically neutral: therefore the ions Ca2+ and Br– combine only in a 1:2 ratio to form calcium bromide, CaBr2. Because no other simpler formula is possible, there is no need to name it “calcium *di*bromide.” CaBr2 can be named using either the Stock method or the older, classic way of naming.

For example, CuCl2 indicates a molecule where one Cu2+ cation associates with two Cl– anions to form a neutral compound. Its systematic name is copper (II) chloride, where copper’s oxidation number is indicated in parentheses. Its older name is cupric chloride.

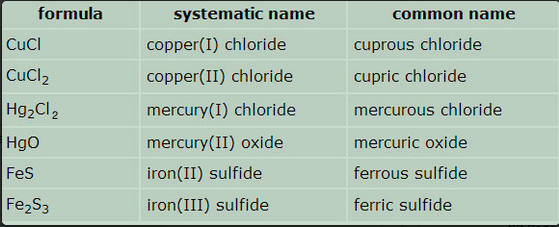
**The Stock Method of Naming**

An ionic compound is named first by its cation and then by its anion. The cation has the same name as its element. For example, K+1 is called  the potassium ion, just as K is called the potassium atom. The anion is named by taking the elemental name, removing the ending, and adding “-ide.” For example, F-1 is called fluoride, for the elemental name, fluorine. The “-ine” was removed and replaced with “-ide.” To name a compound, the cation name and the anion named are added together. For example, NaF is also known as sodium fluoride.

If either the cation or the anion was a polyatomic ion, the polyatomic ion name is used in the name of the overall compound. The polyatomic ion name stays the same. For example, Ca(NO3)2 is called calcium nitrate.

For cations that take on multiple charges (typically transition metals), the charge is written using Roman numerals in parentheses immediately following the element name. For example, Cu(NO3)2 is copper (II) nitrate, because the charge of two nitrate ions (NO3−1) is 2(-1) = -2. Since the net charge of the ionic compound must be zero, the Cu ion has a 2+ charge. This compound is therefore, copper (II) nitrate. The Roman numerals in fact show the oxidation number, but in simple ionic compounds this will always be the same as the metal’s ionic charge.

**The Old, Classic, or Common Way of Naming**



**Names of some ionic compounds**: Common, or trivial, names of compounds are sometimes used in informal conversations between chemists, especially older chemists. Systematic names are formal names that are always used in print.

Since some metallic elements form cations that have different positive charges, the names of ionic compounds derived from these elements must contain some indication of the cation charge. The older method uses the suffixes -ous and -ic to denote the lower and higher charges, respectively. In the cases of iron and copper, the Latin names of the elements are used (ferrous/ferric, cuprous/cupric). This system is still used, although it has been officially supplanted by the more precise, if slightly cumbersome, Stock system. In both systems, the name of the anion ends in -ide.

**Naming Molecular Compounds**

Molecular compounds are named using a systematic approach of prefixes to indicate the number of each element present in the compound.

**Key Points**

* In nomenclature of simple molecular compounds, the more electropositive atom is written first and the more electronegative element is written last with an **-ide suffix.**
* The Greek **prefixes** are used to dictate the number of a given element present in a molecular compound.
* Prefixes can be shortened when the ending vowel of the prefix “conflicts” with a starting vowel in the compound.
* Common exceptions exist for naming molecular compounds, where trivial or common names are used instead of systematic names, such as ammonia (NH3) instead of nitrogen trihydride or water (H2O) instead of dihydrogen monooxide.

**Key Terms**

* **nomenclature**: A set of rules used for forming the names or terms in a particular field of arts or sciences.
* **electronegative**: Tending to attract electrons within a chemical bond.
* **electropositive**: Tending to not attract electrons (repel) within a chemical bond.

**Chemical Nomenclature**

The primary function of chemical nomenclature is to ensure that a spoken or written chemical name leaves no ambiguity concerning to what chemical compound the name refers. Each chemical name should refer to a single substance. Today, scientists often refer to chemicals by their common names: for example, water is not often called dihydrogen oxide. However, it is important to be able to recognize and name all chemicals in a standardized way. The most widely accepted format for nomenclature has been established by IUPAC.

Molecular compounds are made when two or more elements share electrons in a covalent bond to connect the elements. Typically, non-metals tend to share electrons, make covalent bonds, and thus, form molecular compounds.

**Rules for Naming Molecular Compounds:**

1. Remove the ending of the second element, and add “ide” just like in ionic compounds.
2. When naming molecular compounds prefixes are used to dictate the number of a given element present in the compound. ” mono-” indicates one, “di-” indicates two, “tri-” is three, “tetra-” is four, “penta-” is five, and “hexa-” is six, “hepta-” is seven, “octo-” is eight, “nona-” is nine, and “deca” is ten.
3. If there is only one of the first element, you can drop the prefix. For example, CO is carbon monoxide, not monocarbon monoxide.
4. If there are two vowels in a row that sound the same once the prefix is added (they “conflict”), the extra vowel on the end of the prefix is removed. For example, one oxygen would be monooxide, but instead it’s monoxide. The extra o is dropped.

Generally, the more electropositive atom is written first, followed by the more electronegative atom with an appropriate suffix. For example, H2O (water) can be called dihydrogen monoxide (though it’s not usually). Organic molecules (molecules made of C and H along with other elements) do not follow this rule.

**Examples of Molecular Compound Names:**

* SO2 is called sulfur dioxide
* SiI4 is called silicon tetraiodide
* SF6 is called sulfur hexafluoride
* CS2 is called carbon disulfide

**Naming Acids and Bases**

Acid names are based on the anion they form when dissolved in water; base names follow the rules for ionic, organic, or molecular compounds.

**Key Points**

* Acids are named based on their anion — the ion attached to the hydrogen. In simple binary acids, one ion is attached to hydrogen. Names for such acids consist of the prefix “hydro-“, the first syllable of the anion, and the suffix “-ic”.
* Complex acid compounds have oxygen in them. For an acid with a polyatomic ion, the suffix “-ate” from the ion is replaced with “-ic.”
* Polyatomic ions with one extra oxygen (as compared to the typical polyatomic ion) have the prefix “per-” and the suffix “-ic.”
* Polyatomic ions with one fewer oxygen have the suffix “-ous”; ions with two fewer have the prefix “hypo-” and the suffix “-ous.”
* Strong bases with “-OH” (hydroxide) groups are named like ionic compounds. Weak bases are named like molecular compounds or organic compounds.

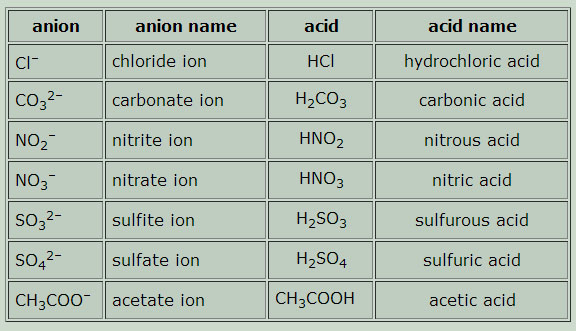
**Key Terms**

* **polyatomic ion**: A charged species (ion) composed of two or more atoms covalently bonded. Also known as a molecular ion.

**Naming Acids**

Acids are named by the anion they form when dissolved in water. Depending on what anion the hydrogen is attached to, acids will have different names.

Simple acids, known as binary acids, have only one anion and one hydrogen. These anions usually have the ending “-ide.” As acids, these compounds are named starting with the prefix “hydro-,” then adding the first syllable of the anion, then the suffix “-ic.” For example, HCl, which is hydrogen and chlorine, is called hydrochloric acid.



More complex acids have oxygen in the compound. There is a simple set of rules for these acids.

1. Any polyatomic ion with the suffix “-ate” uses the suffix “-ic” as an acid. So, HNO3 will be nitric acid.
2. When you have a polyatomic ion with one more oxygen than the “-ate” ion, then your acid will have the prefix “per-” and the suffix “-ic.” For example, the chlorate ion is ClO3–. Therefore, HClO4 is called perchloric acid.
3. With one fewer oxygen than the “-ate” ion, the acid will have the suffix “-ous.” For example, chlorous acid is HClO2.
4. With two fewer oxygen than the “-ate” ion, the prefix will be “hypo-” and the suffix will be “-ous.” For example, instead of bromic acid, HBrO3, we have hypobromous acid, HBrO.

**Naming Bases**

Most strong bases contain hydroxide, a polyatomic ion. Therefore, strong bases are named following the rules for naming ionic compounds. For example, NaOH is sodium hydroxide, KOH is potassium hydroxide, and Ca(OH)2 is calcium hydroxide. Weak bases made of ionic compounds are also named using the ionic naming system. For example, NH4OH is ammonium hydroxide.

Weak bases are also sometimes molecular compounds or organic compounds because they have covalent bonds. Therefore, they are named following the rules for molecular or organic compounds. For example, methyl amine (CH3NH2) is a weak base. Some weak bases have “common” names. For example, NH3 is called ammonia; its name isn’t derived from any naming system.

**Naming Hydrates**

The name of a hydrate follows a set pattern: the name of the ionic compound followed by a numerical prefix and the **suffix -hydrate.**

**Key Points**

* Hydrates are named by the ionic compound followed by a numerical prefix and the suffix “-hydrate. ” The “· nH2O” notation indicates that “n” (described by a Greek prefix) number of loosely bonded water molecules are associated per formula unit of the salt.
* An anhydride is a hydrate that has lost water. A substance that does not contain any water is referred to as **anhydrous.**
* In organic chemistry, a hydrate is a compound of water, or its elements, with another molecule. Glucose, C6H12O6, was originally thought of as a carbohydrate (carbon and water), but this classification does not properly describe its structure and properties.

**Key Terms**

* **hydrate**: A solid compound containing or linked to water molecules.
* **carbohydrate**: A sugar, starch, or cellulose that is a food source of energy for an animal or plant; a saccharide
* **anhydride**: Any compound formally derived from another (or from others) by the loss of a water molecule; a molecule with no water.

**Inorganic Hydrates**

“Hydrate” is a term used in inorganic chemistry and organic chemistry to indicate that a substance contains loosely bonded water. The name of a hydrate follows a set pattern: the name of the ionic compound followed by a numerical prefix and the suffix “-hydrate.” For example, CuSO4 · 5 H2O is “copper(II) sulfate pentahydrate.” The notation of hydrous compound · nH2O, where n is the number of water molecules per formula unit of the salt, is commonly used to show that a salt is hydrated. The “⋅” indicates that the water is loosely bonded to the ionic compound. The prefixes are the same Greek prefixes used in naming molecular compounds

The Greek prefixes used in naming hydrates for numbers 1/2 through 10 are as follows:

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