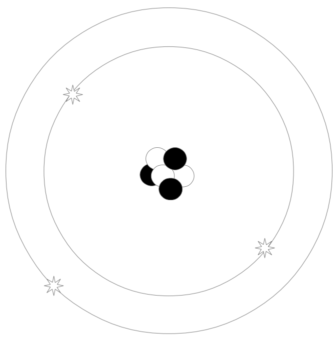
Atoms and Bonding Study Guide

Parts of an atom

All matter in the universe, including stars, buildings, people, and iPods is made of tiny particles called **atoms**. The behavior of atoms is chemistry's main focus.

Inside every atom there are three types of smaller or "subatomic" particle:

- **protons**, found in the nucleus, have a positive electrical charge (+)
- **neutrons**, found in the nucleus, have no electrical charge (0)
- **electrons**, found orbiting the nucleus, have a negative electrical charge (−)



☼ = electron, −

○ = neutron, 0

● = proton, +

Structure of an atom

The protons and neutrons stay close together in the middle of an atom, in an area called the **nucleus**. The electrons fly around the nucleus in one of the atom's **energy levels**, sometimes called **shells**. The lowest level is very small, so it only has room for *two electrons*. The other shells you need to worry about can hold *eight electrons* each. (In a full chemistry course, you'd study the shells in much more detail. This is a simplification.)

An atom will fill up these shells from the bottom up. So for example, if an atom has 11 electrons, it starts by putting 2 in the lowest shell, then 8 in the second shell, and its remaining 1 electron in the third shell.

The top level or shell in an atom is called its **valence shell**, and the electrons in that shell are called **valence electrons**. An atom's number of valence electrons (also called its **valence number**) plays a huge role in how it will react with other atoms. Most chemical reactions end with the involved atoms filling their valence shells to maximum capacity (either 2 or 8 electrons, depending on the atom).

Reading the periodic table

11
Na
22.99
Sodium

Each block on the periodic table represents a different kind of atom known as an **element**. An example is shown at the right. This block contains four pieces of information about the element. Some periodic tables are a little different, but this is what most look like.

At the top is the element's **atomic number**. This tells you how many protons are in the nuclei of this element's atoms. (For a neutral/uncharged atom, the atomic number also tells you how many electrons there are.)

Below this is the element's **symbol**, one to three letters which are used to abbreviate the element's full name. The first letter of a symbol is always an uppercase letter, and the later letters – if it has them – are always lowercase.

Below the symbol is a number with a few digits after the decimal point. This is the **atomic mass**, which tells you the average number of particles in a nucleus of the element. Remember, the nucleus contains both protons and neutrons. Although the number of protons for an element is always the same, the number of neutrons can vary from atom to atom. So the atomic number is always a whole number, but the atomic mass of an element shows a decimal point because it's an average of other numbers. (Sometimes people even call it the *"average atomic mass"* to remind you that it's an average.)

Lastly, at the very bottom of the block is the element's full **name**. The names for lots of elements will be familiar to you, but there are certainly some you've never heard of.

Using the example shown above:

- Atomic number: 11
- Symbol: Na
- Atomic mass: 22.99
- Name: Sodium

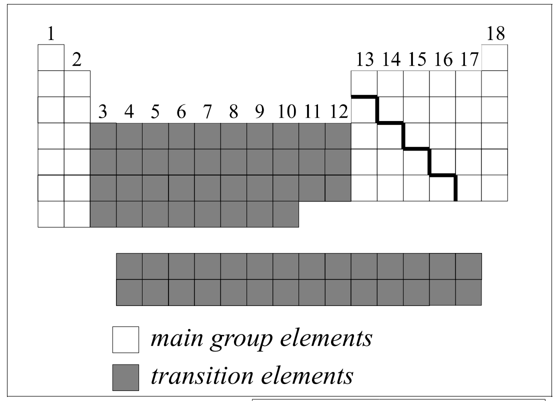
Some periodic tables also include information about the element's **electron configuration**. This is a way of listing the energy levels of every electron in an atom. The easiest way of doing this is to simply say how many electrons are in each level. For sodium, this would be "2-8-1". If you ever take a chemistry class, you may see a more detailed list written in a different style, like this: $1s^2 2s^2 2p^6 3s^1$.

Arrangement of the periodic table

There's other important information on the periodic table which isn't found inside an element's block. This information comes from *where* on the table you find an element.

Most importantly, *elements which are in the same column have similar chemical properties*. That is, all the elements in a single column will probably undergo similar chemical reactions. Each column on the periodic table is referred to as a **group**. There's more than one way of labeling the groups, but the simplest is to just number them from 1 to 18.

The **main group elements** – shown in white at right – are easier to understand than the **transition elements**, which have been shaded. The transition elements all reside in the "skinny" middle section of the table. This guide focuses on main group elements. Transition elements are studied in chemistry classes for grades 11-12 and college students.



The most important thing you can learn about an element from its position on the periodic table is its valence number. Elements in group 1 – the first column of the table – have 1 valence electron. As you count across the eight main groups, the valence number increases by 1 with each step. So, in group 2 there are 2 valence electrons, in group 13 there are 3, in group 14 there are 4, and so on. The only exception to this is helium, found at the top of group 18. Because helium is only the second element on the table, it has only 2 electrons total. So, helium's valence number is 2.

Group	Valence Number
1	1
2	2
13	3
14	4
15	5
16	6
17	7
18	8*

**The element He has valence number 2.*

As stated earlier, an atom's number of valence electrons is the most important thing for determining *how it will react with other atoms*. So, you can see why reading valence numbers off the periodic table can be useful!

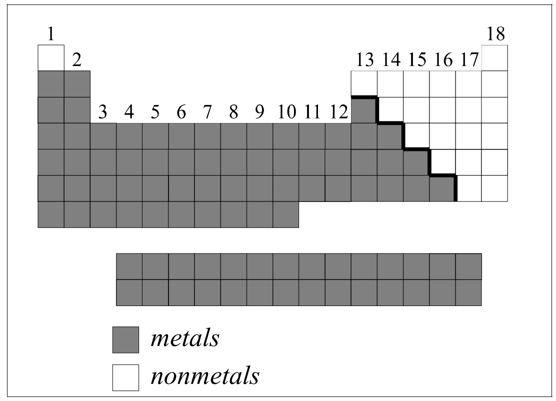
A few of the groups on the periodic table have special names, listed in the table below. You may never be asked, "What is the group number of the alkali metals?" But you probably *will* need to know that some of these groups are very reactive – that is, they will start a chemical reaction with almost anything – and one group is not reactive:

Group	Name	Reactivity
1	**Alkali metals**	High
2	**Alkaline earth metals**	High
17	**Halogens**	High
18	**Noble gases**	Low

Elements in groups 1, 2, and 17 are very reactive because their atoms are so close to having a full valence shell. By adding or subtracting just one or two electrons, they can have their top level filled, so they're eager to take part in reactions. The elements in group 18, on the other hand, already have full valence shells. They almost never get involved in reactions.

If you only remember one thing about the groups on the periodic table, remember that the noble gases – group 18 – don't take part in chemical reactions because they already have full valence shells.

There's one other thing you can easily learn about an element: whether it's a metal or not. The **metals** have a lot in common, like being good conductors of heat and electricity. Usually, metals are shiny and grayish or silvery in color. And most elements are metallic!



On many periodic tables you can find a dark, jagged line through the right side of the table. This line divides the metals, to the left, from the **nonmetals** to the right. The only exception is hydrogen, which is a nonmetal despite its location. Elements along the boundary line on either side are sometimes called **semi-metals** or **metalloids**.

Metals and nonmetals form chemical bonds differently, so being able to tell the difference is important.

Ions and isotopes

Normally, an atom has as many electrons as protons, but that's not always the case. When an atom has missing or extra electrons we call it an **ion**. Since their positively charged protons and negatively charged electrons aren't balanced out, *ions always have an electrical charge*.

The neutrons in an atom are less important, so most elements come in a few "versions" with different numbers of them called **isotopes**. The isotopes of an element still have the same number of protons. Only the neutrons change. Since the isotopes have different numbers of particles in them, their masses are different too. That's why the averaged masses on a periodic table aren't whole numbers!

Ion	An atom where the number of electrons doesn't equal the number of protons. Ions always have an electrical charge.
Isotope	A version of an element with a specific number of neutrons and a specific mass. All isotopes of a given element have the same number of protons.

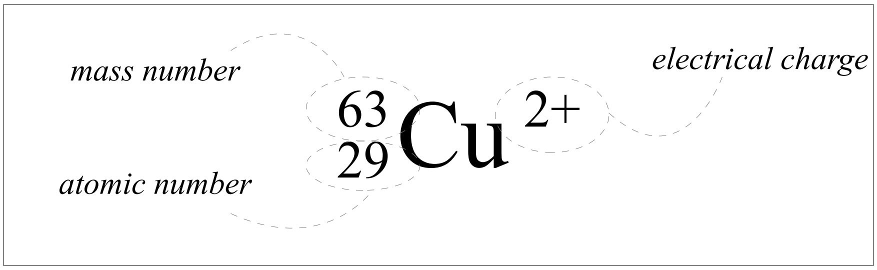
The ions of an element can behave very differently. For example, an oxygen ion with a full valence shell acts more like a noble gas than like normal oxygen. Because being an ion makes such a big difference, you'll never have to wonder whether an atom you're reading about is an ion – you'll be told.

On the other hand, the isotopes of an element are hard to tell apart. The biggest difference is that certain isotopes might be *unstable*: if you wait a while, they give off some radiation and turn into different atoms. Unstable isotopes are said to be **radioactive**. For example, there's an unstable isotope of beryllium with 9 neutrons. It decays into a different beryllium isotope with only 8 neutrons by tossing 1 of its neutrons right out of the nucleus! Some radioactive atoms even decay into atoms of different elements. An example is lead-185 (82 protons and 103 neutrons). Lead-185 decays by releasing two protons and two neutrons in a clump called an alpha particle – changing the lead into a mercury atom!

Number of protons	Determines which element the atom is.
Number of neutrons	Determines which isotope of that element it is. It might be radioactive!
Number of electrons	Determines which ion of that element/isotope it is. And, what it reacts with!

Chemical formulas – Atoms and ions

Information about individual atoms or ions is often shown in a formula that uses the element's symbol and three optional numbers placed around the symbol:



The number in the lower left corner is the atomic number. This number may be omitted because you can easily look up the element's atomic number on a periodic table. If you see a blank space there, don't panic – just check a periodic table.

The number in the upper left corner is the atom's **<u>mass number</u>**. This tells you the total number of protons and neutrons in the atom's nucleus. Since protons and neutrons have almost exactly the same mass, this basically tells you how "heavy" the atom is. It also tells you which isotope you're dealing with. If the mass number is omitted, this means that the exact isotope didn't matter to whoever wrote the formula. (This is often the case.)

In the upper right corner is the electrical charge, if any. Positive charges indicate "missing" electrons and negative charges indicate "extra" electrons. Charges are always written with a number followed by a plus or minus sign, unless the charge is just 1+ or 1−. Then, you can leave off the "1" and write just the plus/minus sign.

Here are some more example formulas:

$^{65}Cu^{2+}$	*Copper isotope with mass of 65. Atomic number omitted. Charge indicates two missing electrons.*
$^{35}_{17}Cl^{-}$	*Chlorine isotope with mass of 35. Atomic number included. Charge indicates one extra electron.*
^{35}Cl	*Chlorine isotope with mass of 35. Atomic number omitted. No charge indicates a neutral atom.*
Cl^{2-}	*Chlorine with mass number and atomic number omitted. Charge indicates two extra electrons.*

You can also talk about a specific isotope without a chemical formula by writing the name or symbol of the element followed by a dash and its mass number. For example: "carbon-14" and "C-14" refer to the carbon isotope with a mass of 14 (6 protons and 8 neutrons). This style is best used only for electrically neutral atoms, since the dash looks like a minus sign. So, don't use it for ions.

How many particles?

By looking at an atom or ion's formula, you can figure out exactly how many of each subatomic particle are inside it. Finding the number of protons is easy: it's the atomic number of the given element. The number of neutrons is the difference between the mass number (which includes both the protons and neutrons) and atomic number. The number of electrons in a neutral atom equals the atomic number. For the number of electrons in an ion, just subtract the ion's charge.

of Protons = Atomic Number
of Neutrons = Mass Number − Atomic Number
of Electrons = Atomic Number − Charge

In the electron formula you must *subtract* the charge even when the charge is negative. For a silicon ion with a charge of 2+, you would take the atomic number (14) and subtract 2, giving you 12 electrons. For a silicon ion with a charge of 2−, however, *you subtract negative two*: $14 - (-2) = 14 + 2 = 16$ electrons. As long as you remember that negative ions have extra electrons and positive ions have missing electrons... you should be able to double-check your work.

When you want to find the number of neutrons in an atom, *you need to know its mass number.* Using the atomic mass from the periodic table instead is not a good substitute because that number is an average. You might end up saying there are 2.4 neutrons, which is impossible... or you might round the atomic mass to a whole number, but you won't know if it's the right whole number for the isotope you're studying. Without having a particular isotope in mind, it's impossible to find the neutron number.

Chemical reactions

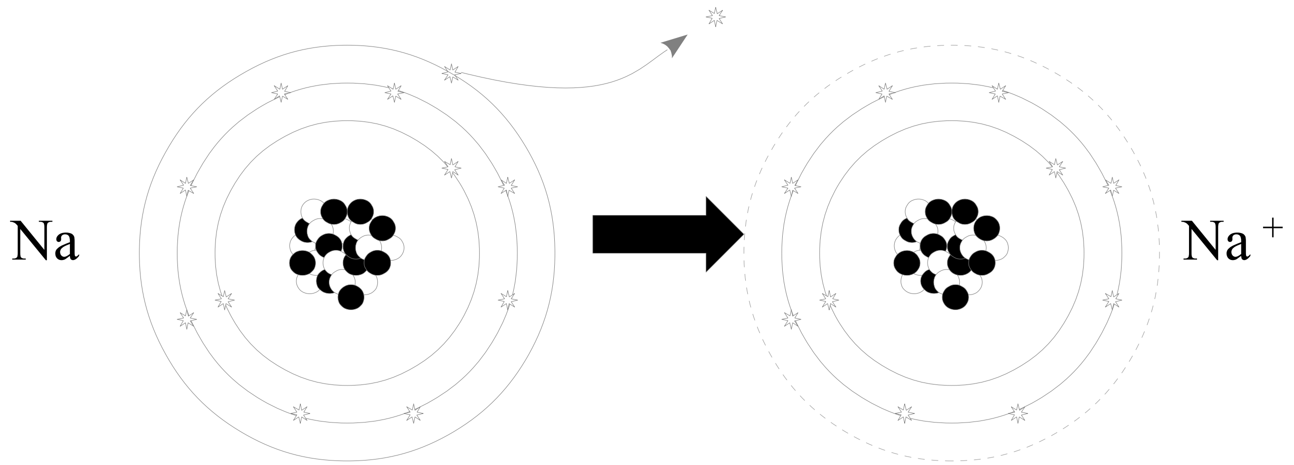
The most stable – or “best” – arrangement of electrons for an atom is when the atom's valence shell is full. Most chemical reactions involve atoms changing from less-stable to more-stable arrangements. Usually this means the atoms end up with 8 valence electrons, unless the atom has only one shell; then it will be full with just 2 electrons. There are two main ways an atom can fill its valence shell:

- *donating electrons or accepting donated electrons* to/from other atoms
- *sharing* electrons with other atoms

In both of these cases, the atoms involved end up very close to each other and become hard to separate. They've formed a **chemical bond** as a result of the donation or sharing of electrons. Once they've bonded, *the atoms form a new substance with different physical and chemical properties than they had before*. Any substance with two or more elements chemically bonded together is called a **compound**.

Ionic bonds

Consider sulfur (S), which has 6 valence electrons: In order to reach a stable arrangement, it needs 2 more electrons. If 2 electrons are given to sulfur from other atoms, its valence shell will be full. On the other hand, the element sodium (Na) has 1 valence electron. To fill up, sodium would need 7 more electrons... but there's an easier way: it could lose the 1 it already has in its valence shell. If that shell is completely empty, it uncovers the shell just below which is already full, as shown below:



A normal sodium atom has 1 valence electron. If that electron is removed from the atom, there are no longer any electrons in the 3rd energy level. The eight electrons in the 2nd level are now valence electrons, so the sodium atom now has a full valence shell.

So, sulfur wants to gain 2 electrons, and sodium wants to lose 1 electron. If two sodium atoms donate their electrons to the same sulfur atom, all three will end up with full, stable shells!

As described earlier, when atoms gain or lose electrons, they become charged and we call them ions. Each of our sodium ions will have a charge of 1+ because they lost 1 electron each. The sulfur ion will have a charge of 2− because it gained 2 electrons. Since they now have opposite charges, the sodium ions will be attracted to the sulfur ion and stay nearby.

Because it involves ions, the attraction of atoms which have gained/lost electrons like this is called an **ionic bond**.

Ionic bonds usually form between atoms of one metallic element and one non-metallic element.

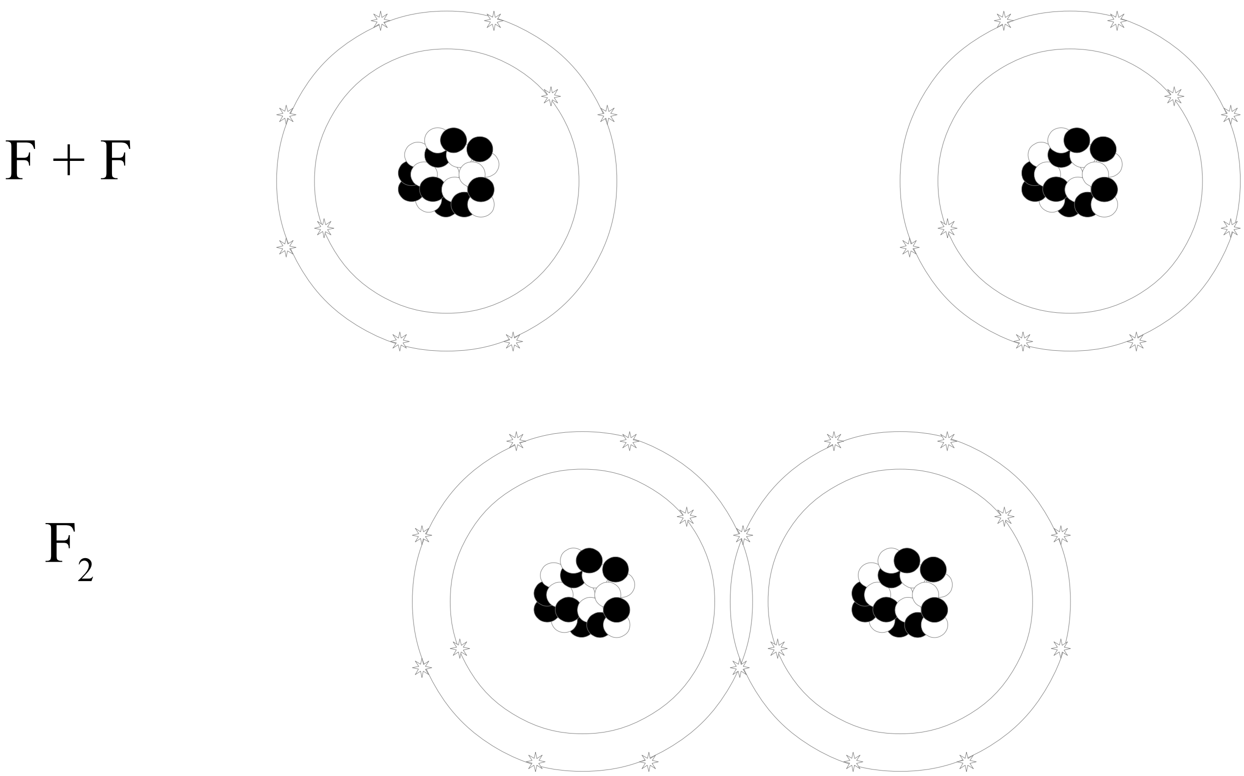
Covalent bonds

Sometimes two atoms will both be a few electrons short of having full valence shells. If there's no simple way for them to give/take electrons, they may instead share a few.

Consider fluorine, which has 7 valence electrons. Two atoms of fluorine can't form a tidy ionic bond. But if a pair of electrons are shared, both atoms can have 8 electrons in their valence shell, as shown in the diagram on the next page.

Because this form of bonding involves a cooperation between two atoms' valence electrons, it is called a **covalent bond**. (Think about the “co-” in words like “coworker” and “coexist” to help you remember this term.) The cluster of atoms in a covalent bond are referred to as a **molecule** as long as they are electrically neutral. If the group has an overall charge like an ion, it's referred to as a **polyatomic ion**.

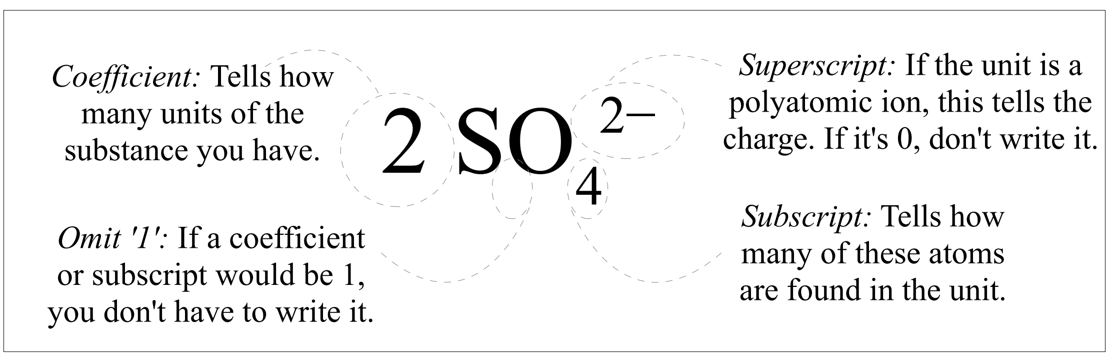
Covalent bonds usually form between non-metallic elements.



A flourine atom has 7 valence electrons. If 2 flourines get close together, they can each share one of their electrons with the other. Since both atoms can use the shared pair, they both get to have a full valence shell.

Chemical formulas – Compounds, molecules and polyatomic ions

Just like the chemical formulas that describe individual atoms or ions, formulas can also describe the chemical makeup of bonded atoms. Because the atoms are bonded together, they can be grouped into clusters called “units” and described with a formula. This example is a polyatomic ion called “sulfate”:



The **superscript** here is just like the superscript on a single atom's formula: it shows the charge of a polyatomic ion. This charge applies to the entire unit, not just the last atom. In the case above, the unit as a whole has a charge of 2− because there are 2 more electrons than protons. Those extra electrons don't belong to the oxygen atoms: they're *shared* by all the atoms in the unit. As you might guess, if the charge is 0, you don't need to write a superscript at all. That means you're looking at a neutral compound instead of a polyatomic ion.

Subscripts written after the symbol for each element tell you how many atoms of that element are found in the unit, but you don't need to write them if that number is '1'. Here, we can see that sulfate contains 1 sulfur atom and 4 oxygen atoms for a total of 5 atoms.

Lastly, in front of the formula is a full-sized number called a **coefficient**. This number tells how many units are present. In the example above, the coefficient shows that you have two sulfate ions. Each one of them contains 5 atoms, so the total number of atoms being described is 2 × 5 = 10. Just like the subscript, you can leave the coefficient off if it would be a '1'.

In an especially complicated formula you may see a set of parentheses. You should think of these just like you would if you saw them in a math problem: they group together certain atoms into a smaller unit, usually a polyatomic ion. As an example, look at the formula for aluminum sulfate:

$$Al_2(SO_4)_3$$

The subscript 3 is written after the parentheses, so it applies to everything inside. That means each unit of this compound has 3 sulfate ions in it. If we're counting atoms, we'll need to multiply all the subscripts inside the parentheses by 3. How many atoms of each element are shown in this formula?

- Al: has a subscript of 2, so there are 2 Al atoms
- S: has no subscript, so there is 1 in each sulfate and therefore 1 × 3, or 3 S atoms overall
- O: has a subscript of 4, so there are 4 in each sulfate and therefore 4 × 3, or 12 O atoms overall

This gives us a total of 17 atoms in each unit of aluminum sulfate: 2 Al, 3 S, and 12 O.

In formulas with a lot of letters, you might have trouble figuring out when one element ends and another begins. For example:

$$H_2PO_2$$

Does this formula show 1 phosphorus (P) and 2 oxygens (O), or does it show 2 poloniums (Po)? If you take a deep breath and read the formula carefully, there's an easy way to tell: Every capital letter is the beginning of a new element's symbol. “PO” means “phosphorus and oxygen” because the O is capitalized. The trickiest examples come from letters that are hard to tell apart, like the lowercase “l” in chlorine (Cl) and the uppercase “I” in iodine (I). Your teachers can have trouble telling them apart, too, so do them a favor and try to write your chemical symbols distinctly!

This summary addresses material typically taught to 9th graders in the United States and is simplified accordingly. It is meant to be representative of the material that appears on many state graduation tests. It is not appropriate as a study guide for a dedicated chemistry course.

This summary was prepared by Steve Stonebraker, a high school science teacher in Ohio, and last revised on 2012-04-05.