Chapter 2 CHEMICAL UNITS AND THEIR IDENTITIES, Part 1

Let's start into chemical units. What are they?

Chemical units are the stuff of your world. A chemical unit can be an atom, a molecule or an ion. The oxygen you need to breathe. The water necessary to all life on the planet. The various minerals in the soil. The DNA of your genes. Every pure substance has a specific chemical identity. This is what defines it. Every property, every characteristic, everything about a substance is determined by who it is chemically. Likewise, all things of the same chemical identity have the same properties. Water is water and it will always be water no matter if it is from the ocean, falls as rain, or is taken from Mars. Water is H_2O. That is the chemical unit. That is its identity.

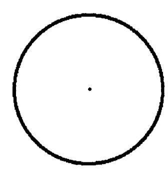
But we're jumping ahead. Let's proceed to the start.

2.1 Atoms

Chemical units begin with atoms. Every atom is made of "subatomic particles", which include protons, neutrons and electrons. (Some of these can be broken down further and there are other subatomic particles, but we won't go there.) Matter is composed of protons, neutrons and electrons but <u>it is the different numbers of these which give us different atoms</u>. Here are nine basics for atoms which you need for right now. Know these.

1. Protons (symbol, p^+) have a 1+ charge.
2. Electrons (symbol, e^-) have a 1− charge.
3. Neutrons (symbol, n^0) have no charge.
4. The size of one single proton or one single neutron or one single electron is incomprehensibly tiny.
5. At the center of the atom lie all of the atom's protons and neutrons. Together, these make up the nucleus. Since all protons and neutrons are incomprehensibly tiny, the nucleus is also incomprehensibly tiny. Too small to even worry about for our purposes.
6. Protons and neutrons have very similar masses. They're not quite exactly the same, but close enough: a neutron is 1.0014 times heavier than a proton. Compared to one proton or one neutron, one electron has very little mass. In fact, the mass of a neutron or proton is about 1800-times the mass of an electron. An electron's mass is at most 0.05% of a whole atom's mass. (Values for the masses are in Section 2.3.)
7. IMPORTANT RESULT FOR MASS: The mass of an atom is almost entirely (99.95% or more) in the nucleus with the protons and neutrons.
8. The electrons are widely spread out in a volume surrounding the nucleus. This volume is actually very large (by many thousands of times) when compared to the size of the nucleus. Although this volume is large compared to the nucleus, "large" to an atom is still very small compared to anything we can normally see.
9. IMPORTANT RESULT FOR SIZE: The size of the atom is almost entirely determined (~99.99%) by the volume in space which is occupied by the electrons.

Your basic picture of an atom is shown at right. OK, the drawing is flat and that's not right: atoms are really three-dimensional, so an atom is more like a ball and not like a circle. The tiny dot in the middle represents the nucleus. The big circle represents the volume occupied by the electrons. Notice that I keep saying "volume occupied by the electrons", but it's not so simple. There's a lot more to that and we will get into that beginning in Chapter 20. For now we'll stick with the simple picture.



Everything about an atom is infinitesimally tiny. The masses of individual atoms start at 0.000000000000000000000002 grams. The sizes of atoms are on the order of 0.0000000001 meters. These are incredibly small numbers.

For the most part, CHEMISTRY INVOLVES ELECTRONS. Not protons. Not neutrons. (Unless you're nuking things.) ELECTRONS DO CHEMISTRY! This will be a recurring theme throughout the entire course. This makes electrons a primary focus in our coverage. We need to know three key pieces of information regarding electrons.

We need to know how many there are.
We need to know their locations within the atom.
We need to know their energies.

Beginning in this Chapter, we will see how many electrons there are. Locations and energies are more complicated; we begin those details in Chapter 20. Although electrons do chemistry, they cannot exist alone in a chemical unit. Due to their charges, their numbers are related to the numbers of protons, as we shall see. In fact, it is the charge attraction between protons and electrons which gives us atoms. Thus, protons are a necessary part of the picture. For most cases, neutrons influence chemistry very little. They're mostly dead-weight, but they are a very important part of the weight. So why are they there? They're needed to hold protons together in the nucleus. Since positive charges repel each other, protons are mutually repulsive toward one another; neutrons keep them all together. They've confirmed atoms with up to 112 protons, along with 114 and 116; others up to 118 have been claimed but not confirmed. (They have to be confirmed before they're considered official.)

Here's one more thing about this charge business: a neutral atom is defined as having no charge. If an atom is neutral, then the number of protons and the number of electrons must be the same and the pluses and the minuses cancel out to zero. But life is not always neutral, as we shall see.

2.2 Identity

So you have an atom. An atom of what? Who is it? What is its identity?

It takes only one thing to identify an atom: the number of protons in its nucleus. Once you know this, you know who it is. This is the only criterion of identity. Thus, this is a very important number, and it gets the name "atomic number" and the symbol Z. (Why Z? It's from the German Zahl for number.)

$$\text{atomic number} = Z = \text{number of protons}$$

The different atomic numbers give us different identities for atoms. Each identity is an "element". An element is an atom of a specific identity or a number of atoms of the same identity. It's also a name. It's who they are. Like hydrogen. Or helium. Or oxygen. Or gold. Or praseodymium. (Praseodymium? Hey, I didn't come up with the names. But it could be worse. There could've been an element named vitameatavegaminium.) The confirmed atoms with up to 112 protons have been named. The numbers and the names of those elements are on the inside back cover of this book. Look over some of these. Some names are simple. Some are not. Many names you've probably heard of. (At the time of this printing, the atoms with 114 and 116 protons were not yet named.)

You are known by your name but sometimes you just use your initials. A similar relationship applies for element names. For many, we use their initials or another abbreviation. These are called symbols. All symbols are either a single capital letter or one capital letter followed by one lower case letter. (Elements beyond $Z = 112$ have three letters as a temporary symbol.) Most symbols correspond closely to the name. "H" stands for hydrogen. "P" stands for phosphorus. "Ne" stands for neon. "Al" stands for aluminum. "Mg" stands for magnesium. Notice with Mg that the second letter of the symbol does not have to be the second letter of the name. Some of the element symbols are not even close to their names. Actually, they're not close to their English names. For these, the symbol was derived from the name in another language, most commonly Latin. "Ag" stands for silver but the symbol comes from argentum (Latin). (People who do silver work are sometimes called argentists.) "Pb" stands for lead but the symbol comes from plumbum (Latin). (The word plumbing was derived from this. Lead was used for plumbing way back since Roman times, long before it was known to be so deadly.) "K" stands for potassium but the symbol comes from kalium (Latin). (You need the right amount of potassium in your bloodstream. If you have too little, the clinical condition is called hypokalemia, from hypo- and kalium.) Not all of the oddball names are from Latin. "W" stands for tungsten, but the symbol comes from wolfram (German).

The symbols are also in the element list on the inside back cover. Look over them. Become familiar with some. Actually, you should memorize some of the symbols and names since they are such an important part of the language (verbal and written) of chemistry. Which ones should you memorize? Ask your instructor.

OK, that ends it for element identity: it's the number of protons, atomic number, Z. Now let's talk about neutrons. Earlier I said that they were mostly dead-weight. So why bother with them? It's because mass is an extremely important part of how we actually do things, and neutrons are a large part of an atom's mass. We have no choice. Although they do little, we must be aware of their presence and their contribution to mass. As noted above, the mass of an atom is almost entirely due to the protons and neutrons, so we need to know how many neutrons there are for mass reasons. This gives us a new term, the "mass number", which gets the symbol A (from German for Atomgewichte). It is defined as the sum of the number of protons and number of neutrons.

mass number = A = number of protons + number of neutrons

Notice that this is not a true mass. In fact, it is simply a count of total nuclear particles.

Let me illustrate atomic number and mass number with a few examples.

A. An atom of 7 p^+ and 7 n^0 has atomic number 7, which means it's a nitrogen (N) atom. The mass number is 14.

B. An atom of 17 p^+ and 18 n^0 has atomic number 17, which means it's a chlorine (Cl) atom. The mass number is 35.

C. An atom of 42 p^+ and 53 n^0 has atomic number 42, which means it's a molybdenum (Mo) atom. The mass number is 95.

D. An atom of 42 p^+ and 57 n^0 has the same atomic number 42, which means it's also a molybdenum (Mo) atom. The mass number is now 99.

E. An atom of 43 p^+ and 56 n^0 has atomic number 43, which means it's a technetium (Tc) atom. The mass number is also 99.

(Most people never hear of technetium. It's radioactive. If you ever go to a hospital for a "radio-diagnostic" procedure, the chances are good that they'll inject you with a technetium compound. The "radio-" prefix in "radiodiagnostic" comes from radioactive, not from FM or AM.)

Remember that only the number of protons is element identity. Examples C and D above are the same element but with different numbers of neutrons. They'll have different masses but they are the same element. Examples D and E have the same mass number, but they have different element identities.

Variations in neutron numbers give rise to "isotopes", which are atoms of an element with specific numbers of neutrons. Examples C and D represent two different isotopes of molybdenum, one with 53 n^0 and one with 57 n^0. Nature gave us several isotopes for most of the elements on this planet. Chlorine has two. Hydrogen has three. Calcium has six. Molybdenum has seven. Only 22 of the elements on this planet are found naturally in only one isotope; fluorine, phosphorus, sodium and gold are examples.

Sometimes it is important that we distinguish one particular isotope for a particular element. The symbolism is then modified to incorporate this: immediately before the usual element symbol, the mass number is placed as a superscript. The mass number is also used when we speak the isotope's name. Let's see how this is done for the five cases above.

A. The specific isotope symbol is ^{14}N. We say "nitrogen-14".
B. The specific isotope symbol is ^{35}Cl. We say "chlorine-35".
C. The specific isotope symbol is ^{95}Mo. We say "molybdenum-95".
D. The specific isotope symbol is ^{99}Mo. We say "molybdenum-99".
E. The specific isotope symbol is ^{99}Tc. We say "technetium-99".

There is another version which also shows atomic number as a subscript right below the superscript mass number. An example is $^{14}_{7}N$. This is not as common as just the ^{14}N version since the atomic number 7 automatically means N and N automatically means atomic number 7. But this method is used in some applications.

That wraps up mass number. Now we need to start talking about real masses.

2.3 Mass

Mass is extremely important to the historical development of chemistry and to doing chemistry even today. Ultimately, the number of atoms in a given sample can be related to the mass of the sample. We do that relationship later. Right now, we focus on the masses of individual atoms. This is fairly straightforward except for two minor details: atoms have extremely small masses by themselves which makes direct measurement impossible; and, we have to account for isotopes since different isotopes of the same element have different masses.

The mass of any single atom is incredibly tiny, in the range of 10^{-22} to 10^{-24} g. These are not convenient numbers to work with, so we often use an "atomic mass unit", u, for masses in the range of individual atoms. One u is an incredibly small mass, 1.6605×10^{-24} g, and this is close to the mass of a single proton or neutron: 1.0073 u for a proton and 1.0087 u for a neutron. (In comparison, an electron mass is only 0.00055 u.) Let's look at some examples of actual masses for atoms.

One atom of ^{19}F has an actual mass of 18.9984032 u, which is close to its mass number, 19.
One atom of ^{35}Cl has an actual mass of 34.96885271 u, which is close to its mass number, 35.
One atom of ^{37}Cl has an actual mass of 36.96590260 u, which is close to its mass number, 37.
One atom of ^{197}Au has an actual mass of 196.966569 u, which is close to its mass number, 197.

Since electrons contribute very little mass relative to protons and neutrons, the true mass is very close to the mass number. In most cases, the true mass is a bit less than the mass number due to another factor, called mass defect, but I won't go into that right now.

Another unit for the atomic mass unit is "dalton", which goes by the symbol Da. It's just an alternate term. If you see it somewhere, just remember Da = u.

Remember that these are masses for individual atoms of a specific isotope. For a real sample of zillions of atoms, we have to account for the different isotopes which may be present. In these cases, we work with an average. I will illustrate this for chlorine, which comes in the two isotopes cited above, ^{35}Cl and ^{37}Cl.

Nature mixed the two isotopes ^{35}Cl and ^{37}Cl in all chlorine compounds on Earth. They're not in equal amounts, however: Nature put more ^{35}Cl on Earth than ^{37}Cl. How much more? 75.78% of all chlorine atoms are ^{35}Cl and 24.22% are ^{37}Cl. These numbers are called percent abundance. They are also called natural abundance, since they are <u>the percent abundance as found naturally on Earth</u>. On some other planet, the numbers can be different. Humans can also process certain isotopes and cause the amounts to change. Except for those cases, the natural abundance will apply regardless of whatever ocean, whatever mine or wherever on Earth the sample was taken.

Given these percent abundances, if we have 100.00 zillion chlorine atoms, then we have 75.78 zillion ^{35}Cl atoms and 24.22 zillion ^{37}Cl atoms. In practical work dealing with mass, we work with the average mass, which takes into account the abundance for each:

$$\frac{(75.78 \text{ zillion } ^{35}\text{Cl weighing } 34.96885271 \text{ u}) + (24.22 \text{ zillion } ^{37}\text{Cl weighing } 36.96590260 \text{ u})}{100.00 \text{ zillion total}}$$

This is the same as taking the weighted average, using the natural abundances written as a decimal instead of as a percent.

$$\begin{gathered}(0.7578)(34.96885271 \text{ u}) + (0.2422)(36.96590260 \text{ u}) = \\ 26.50 \text{ u} + 8.953 \text{ u} = \\ 35.45 \text{ u}\end{gathered}$$

The answer is 35.45 u either way. This is the <u>average</u> mass of Cl atoms, taking into account the natural isotopes. This value is called the "atomic mass". A related term is "atomic weight" which is specifically a <u>relative</u> atomic mass and it does not carry a unit. Thus, the atomic weight of Cl is 35.45. Notice that the atomic mass and the atomic weight are the same number but atomic mass carries the unit u. That may seem like a trivial distinction, but that is how the terms are set up.

Based on our calculation above, the atomic mass of chlorine is 35.45 u. But be careful! There is no single atom of chlorine which has a mass of 35.45 u! Not on earth. Not anywhere. It's an average. This manner of calculation is general for any element with two or more isotopes in order to determine the atomic mass. For each isotope, you multiply the natural abundance (as a decimal number) times the mass of that isotope; then you add these up for all isotopes.

If, however, the element comes in only one natural isotope, then the atomic mass is that single isotope's mass. For example, Nature made all fluorine atoms on earth as ^{19}F, so the atomic mass is the same as that isotope's mass, 18.9984032 u. Nature made all gold atoms on this planet as ^{197}Au, so the atomic mass is the same as that isotope's mass, 196.966569 u. There are no calculations for these cases.

Some of the atomic masses are known to nine sigfigs and some are only known to four. We will work with four for the most part in this book. Feel free to use more if you wish (or are required). I'll warn you: atomic masses are routinely evaluated by an international committee and revisions are frequently done. If you have an older table of values, some of the values may be wrong.

2.4 The elements and the Periodic Table

I'm going to refer back to this section in a later Chapter. Put a red star in the margin.

Pure substances can be characterized as elements or compounds. An element can exist by itself, such as a piece of iron metal or aluminum foil. We refer to this as an "elemental form" of an element.

An element can also be joined in a chemical unit with other elements to form a "compound". Every compound has its own chemical identity and its own set of properties. We'll talk gobs about compounds later. Right now I want to spend a bit of time talking about some of the elements, their elemental forms and what compounds they occur in. First, let's talk about the Periodic Table.

The Periodic Table is a tabular presentation of all the elements. It's an incredibly important tool in chemistry. There's one on the inside front cover of this book. Take a look at it. Get used to seeing and using a Periodic Table because we'll be doing a lot with it. In the Periodic Table, the elements are located in boxes in left-to-right sequence of atomic number. The contents of the boxes typically include the atomic number at the top of each box, the element symbol in the center, and the atomic weight at the bottom. For example, look in box #17: 17 is the atomic number (= number of protons), Cl is the symbol (chlorine), and the atomic weight is 35.45, just as we calculated above.

Although most boxes have the atomic weight of the element at the bottom, the boxes for most of the radioactive elements have a mass number for only one isotope instead, usually in parentheses. Why? Some of these elements are not even found naturally and are instead produced by artificial means. Others may have a natural abundance but the isotope distribution varies due to the radioactivity. Only three radioactive elements have a consistent isotope distribution, so they do show their atomic weight. These three are Th (#90), Pa (#91) and U (#92).

The Periodic Table is called "periodic" because the elements are ordered according to periodic patterns of electron arrangements. Since ELECTRONS DO CHEMISTRY, this arrangement also gives periodic patterns of chemical reactivity. As a consequence of these patterns, the Periodic Table has a specific array of vertical columns and horizontal rows and this gives us the odd shape overall. In Chapter 22, we'll explain this overall pattern. But that's a ways off from now.

The 18 vertical columns in the large portion are called "Groups". These are numbered 1 - 18 and we refer to a specific Group by that number, such as Group 13. Older Periodic Tables used one of two different A/B designations with numbers running 1 - 8 or 0 - 8; that was confusing, so they went to a straight 1 - 18 system. You'll still find Periodic Tables with the old numbers, so be careful with those.

The Groups are collectively separated into the "Main Groups" and the "Transition Metal Groups" (or "Transition Element Groups"). There are eight Main Groups. These include the two columns at the far left (Groups 1 and 2) and the six columns at the far right (Groups 13 - 18). Every other Group is a transition element Group (or transition metal Group). This distinction is of historical origin and it remains very useful since the chemistry of Main Group elements can be quite different from the chemistry of the transition elements.

Now consider the horizontal rows. The horizontal rows are called "Periods". There are seven. They are numbered 1 - 7, with 1 at the top and 7 at the bottom. These are often referred to as First Period, Second Period, etc.

The combination of Group and Period provides an address. For example, selenium (Se, #34) is in the Fourth Period of Group 16. Cesium (Cs, #55) is in Group 1, Sixth Period. The importance of a specific Group and Period will be discussed in greater detail in later Chapters.

Periodic Tables are usually printed with a 2 × 14 chunk of elements at the very bottom. These actually belong in the sixth and seventh rows between elements 57/72 and between elements 89/104. Just look at how the atomic numbers run. This 2 × 14 chunk appears separately at the bottom for printing reasons, not for chemistry reasons. If that chunk was in the right spot, the Table would be too wide to print well on a typical page. Those elements haven't been abandoned; they belong with the others although they do not get any Group number at all. They are also transition elements, although sometimes they are further called "inner transition elements". You can occasionally find printings of the Periodic Table with that chunk squeezed in with the others. This printing format is called the "long form". Believe me, it's really long.

Overall, most (~4/5) of the elements are metals in their elemental forms. This leads to another categorization of the elements as metals or otherwise. "Otherwise" can be further broken into two types, metalloids and nonmetals.

Many of the metal elements are very common to you in their elemental forms. Iron (Fe, #26) in steel. Aluminum (Al, #13) foil. Gold (Au, #79), platinum (Pt, #78) and silver (Ag, #47) in jewelry. Copper (Cu, #29) wire. Tungsten (W, #74) filaments in incandescent bulbs; mercury (Hg, #80) in fluorescent bulbs. Tin (Sn, #50) coatings in some food cans. Etc. While the classification of an element as a metal is fairly common, the distinction between nonmetal and metalloid is not so obvious.

"Nonmetal" is somewhat straightforward: an element is a nonmetal if it has no significant parallel to the general properties of metal elements. There are 18 nonmetal elements; these include the common elements hydrogen (H, #1), helium (He, #2), nitrogen (N, #7), oxygen (O, #8), phosphorus (P, #15), etc. All of these are near the right side of the Periodic Table except for H which is way out in left field. "Metalloid" is intermediate: these elements have some properties which are similar to metals and some which are not. There are only six metalloids: boron (B, #5), silicon (Si, #14), germanium (Ge, #32), arsenic (As, #33), antimony (Sb, #51) and tellurium (Te, #52). These are located in the Periodic Table roughly diagonal between the metal and nonmetal elements, although they only go down to the Fifth Period.

Your instructor may require you to memorize which elements are nonmetals and which are metalloids. (Everything else is a metal.) Or you may only need to memorize which are metals and which are not metals, without worrying about the nonmetal and metalloid distinction. Simply distinguishing metals from "not metals" proves very useful in applications later on, so let me show you how to distinguish at least this much: think of "PoBiSnAl". PoBiSnAl includes the four elements Po, Bi, Sn and Al (polonium, bismuth, tin and aluminum). Find these on the Periodic Table (#84, #83, #50 and #13). All four of these elements are metals, and so also are the elements directly below them (Ga, In, Tl and Pb). All together, these elements are as far as the metal elements go to the right in the Periodic Table. Except for oddball hydrogen (H), <u>all elements</u> in <u>all Groups</u> to the left of these are also metal elements. Thus, if you just remember PoBiSnAl, then you can locate and distinguish metals from "not metals" in the Periodic Table. Try it. Don't forget the 2 × 14 chunk down below; they belong in the metals, too. The extent of the metals in the Seventh Period is not known for sure since little is known about many of those elements. Nevertheless, the range of metals should extend to #116 or so.

Pronouncing "PoBiSnAl" sounds something like Pepto-Bismol, the common remedy for upset stomachs and nausea. This can help you remember it. There's also a chemical connection: Pepto-Bismol uses a bismuth compound as its active ingredient. Read it on the label sometime.

2.5 Group tours

Let's tour the Periodic Table by Group or by collection of Groups. Some Groups have specific (and historical) names and we'll bring these in also.

Group 1 includes hydrogen (H) and the "alkali metals", lithium (Li) down through francium (Fr). Sodium (Na) and potassium (K) are the most common alkali metals. You won't usually find the elemental forms for these but they do look like a typical metal. They're also soft, so soft that you can cut them with a knife. The reason that you won't usually find the elemental forms is that they're extremely reactive. They can explode just with water. Although the elemental forms are not common, compounds of sodium and potassium are extremely common and can be found just about everywhere on Earth. Compounds of both are necessary for all life on the planet. (I mentioned your potassium balance earlier.) Lithium (Li) and its compounds are less common, but they do find important use in some medicines and in lightweight batteries. The lowest members of the Group are not very common.

Group 2 is called the "alkaline earth metals". Unfortunately, this name is close to alkali metals of Group 1, so be sure to remember the difference. Group 2 is topped by beryllium (Be) which happens to be very toxic as a metal and in most of its compounds. In fact, Be is considered the most toxic of all nonradioactive metals. But not all of its compounds are deadly: some of you wear a fairly safe beryllium compound, namely emeralds. Below Be are magnesium (Mg) and calcium (Ca) which are very common in many compounds all over the planet. Compounds of both are essential for life. Calcium compounds, for example, make up marble, gypsum, teeth, bones, etc. The lower elements in the Group are not very common unless you've had a "barium swallow". A barium swallow is a diagnostic procedure in which you swallow a barium compound and X-rays are taken. Barium blocks X-rays well, so they get a good picture of your digestive tract.

After Group 2, the Periodic Table dips down to the Transition Metal Groups 3 - 12, topped by the row of scandium (Sc) through zinc (Zn). All of these have a very wide range of properties. They also have a very wide range of occurrence and use, both in their elemental forms and in their compounds. Many of these are familiar as normal metals by themselves or combined with other metals in alloys. I already mentioned iron (Fe), gold (Au), copper (Cu), tungsten (W), mercury (Hg), platinum (Pt) and silver (Ag). Mercury is the only liquid metal element at typical room temperature. (Some other metals, however, melt close to room temperature. For example, cesium (Cs, Group 1) and gallium (Ga, Group 13) melt below body temperature.) Many of the transition elements are also critically important biologically, even some of the unusual ones. Of the ten transition elements in the Fourth Period, the eight from vanadium (V) through zinc (Zn) form compounds which Nature uses in different ways in different organisms. Iron is the most familiar since it forms hemoglobin in our blood, but it also forms a bunch of other enzymes and proteins which are crucial to survival at all levels of life including humans. Many people take iron supplements for better health. What many people do not realize, however, is that iron is toxic. For example, thousands of children are poisoned each year by getting into their parents' iron supplements and eating them. Some of these children die. Iron is toxic to adults, also, and anyone can die from too much iron. Humans simply do not tolerate iron overdose well. Yes, you need iron to live but too much can kill. In fact, <u>everything you need to live is toxic</u> at some amount. This is one of Nature's greatest ironies. You should remember this.

Iron and other metals used in biology are sometimes called "biometals". Let's consider some less familiar examples. Cobalt (Co) is another one essential to humans; for example, vitamin B_{12} is a cobalt compound. Manganese (Mn) compounds are critical to photosynthesis. (Actually, manganese and magnesium (Mg, Group 2) compounds are involved in photosynthesis. Chlorophyll is a magnesium compound, but manganese compounds are necessary in later stages of photosynthesis.) Zinc (Zn) compounds are very widespread at all levels of life in a number of proteins and enzymes. So also are copper (Cu) compounds. In fact, copper compounds are used instead of iron compounds to transport oxygen in the blood of some critters; many of those have green-blue blood instead of red due to the color of the copper compound. There are enzymes which are nickel (Ni) compounds. Vanadium (V) compounds pop up periodically in biology. Sea squirts actually use quite a bit of vanadium compounds. What are sea squirts? Strange, that's what they are.

These transition elements cited so far are just from the Fourth Period. Biometals go further down also. The enzyme responsible for biological nitrogen fixation is a compound of iron and molybdenum (Mo), along with other nonmetal elements. Biological nitrogen fixation is the process whereby elemental nitrogen from the air is converted into ammonia, a nitrogen compound; this is a critical step in the biological nitrogen cycles on Earth. Interestingly, although molybdenum is not one of the more familiar elements to humans, it is very familiar to Nature and Nature uses molybdenum compounds in many organisms. This includes humans: you need molybdenum to live and, without enough of it, you die. Below molybdenum lies tungsten (W). In recent years they have discovered that tungsten compounds are used by Nature in "hot" applications. This includes microbes which live near geysers where temperatures far exceed anything which a typical organism could survive.

Speaking of biological roles, humans have also tapped into transition element compounds for medicinal applications. Platinum (Pt) compounds are among the most widely used anticancer drugs. Gold (Au) compounds have been used to treat arthritis for many years. Certain vanadium (V) compounds mimic the role of insulin and these have been studied for use in diabetes. I mentioned earlier that technetium (Tc) compounds are used in radiodiagnostics. Mercury (Hg) compounds were once used widely to treat a variety of ailments but this has decreased tremendously. All of these metals in medicine have toxic side effects, however, and their use must be monitored.

That's enough of the transition elements for now. Let's mosey on to other parts of the Periodic Table. We'll go downstairs to the inner transition elements.

Actually, there's not so much to talk about here in terms of typical encounters. The elements of the stretch from cerium (Ce) through lutetium (Lu) are not very common. There is one element in there whose compounds are important in medical use. Gadolinium (Gd) compounds are routinely used in medicine for MRI (Magnetic Resonance Imaging). If you've ever had MRI, you may have had a gadolinium compound injected into you. Why gadolinium? Ask me in Chapter 24. MRI is used a lot these days. Very good method. Nifty pictures.

The lower stretch of the inner transition elements includes thorium (Th) through lawrencium (Lr). All are radioactive and many of their major uses are nuclear things. Uranium (U) is here, and it is perhaps the one element people associate the most with radioactivity. Ironically, its radioactivity is extremely slow and it does have some non-nuclear uses such as forming very hard alloys with other metals. One other element in this stretch can be found in many homes. Americium (Am) is used in battery-operated smoke detectors. Yes, you may have a nuclear device right in your own home. The americium does its nuke thing in the sensor chamber to detect the smoke. It's ^{241}Am. Next time you change the batteries (at least once or even twice a year AND DON'T FORGET!), you may notice the nuclear warning labels inside but don't go messing with it.

OK, let's get back to the main portion of the Periodic Table. Group 13. Aluminum (Al) is the most familiar of the Group and its elemental form is a household staple as foil and some cookery. Just above is boron (B) whose compounds are essential to many forms of life, especially plants. Lower members in the Group are much less encountered.

Proceeding to Group 14, we first find carbon (C). Its compounds are ubiquitous in Nature. Sugars. Proteins. DNA. All are carbon compounds. So also is the carbon dioxide of the air and the carbonate minerals of the earth. Even the elemental forms of carbon are common. The most common are graphite and diamonds. These two are very different, but pure graphite and pure diamond are elemental forms of carbon. Below carbon lies silicon (Si), a major component of the compounds found in the earth's crust as silicates. Elemental silicon is used massively in the electronics industry. The two lowest members of the Group, tin (Sn) and lead (Pb), are metals. Lead compounds are an important environmental concern.

Group 15 is called the "pnictogens", although this term is not as common as some of the others. (The "p" is silent, as in the word pneumonia.) Nitrogen (N) is very widespread in its elemental form and in its compounds. In fact, nitrogen is one of the most familiar elemental forms, although you may not know it. It's ~78% (on a dry basis) of the atmosphere you breathe. Nitrogen compounds are in all life forms on the planet and are absolutely essential. This is where biological nitrogen fixation (mentioned above) comes in. Below nitrogen lies phosphorus (P) whose compounds are also absolutely essential to all life forms on Earth. Phosphorus forms many phosphate compounds which are common minerals, including your teeth. The lower pnictogens are less common. Arsenic (As) compounds are very toxic in general and they are a serious environmental concern in some parts of the world; nevertheless, arsenic compounds are used by some biological organisms and a few arsenic compounds are used to treat cancer. A medicinal role for bismuth (Bi) compounds was mentioned earlier (Pepto-Bismol); roles for other bismuth compounds continue to be explored in recent years.

Group 16 is called the "chalcogens" and again the top two members are the most common and most important. Animals need elemental oxygen (O) to live; the atmosphere is ~21% of this gas. But this is another great example of Nature's irony: although it is essential to breathe oxygen gas, it is toxic in high concentrations. Fortunately, those toxic concentrations are not readily encountered. Scuba divers can encounter this, however, and there are grave restrictions on diving with pure oxygen. Even the normal 21% of the atmosphere can cause health problems over one's lifetime. Many people take "anti-oxidants" to protect themselves from the toxicity of the oxygen they need to live. (Oxygen is the ultimate "oxidant" for humans.) Elemental oxygen can cause cancer, too, but don't tell the government or they'll want to ban it from the workplace.

Below oxygen in Group 16 is sulfur (S), which some people are familiar with as a yellow powder in its common elemental form. Sulfur compounds are also widespread in Nature and they occur in all life forms. Many proteins and enzymes are sulfur compounds. In addition to these general and widespread uses, Nature also has specific uses in specific organisms. The essences of garlic and of onion (including what makes you cry) are sulfur compounds. The smell of a skunk is due to sulfur compounds. Although some sulfur compounds are among the most foul-smelling of all substances, not all sulfur compounds smell: the artificial sweetener, saccharin, is also a sulfur compound. The lower elements of Group 16 are not nearly as common as oxygen and sulfur. Curiously, although selenium (Se) compounds tend to be very toxic, once again Nature throws a curve ball: selenium is an essential element for humans because we need it for certain enzymes. In small amounts, selenium can kill you; in very tiny amounts, you need it to live. Other selenium enzymes are found in some simpler life forms.

Continuing our trek, we hit Group 17, the "halogens". The halogens are led by fluorine (F) whose elemental form is the most aggressive, the most reactive, the most violent of all elemental forms. Actually, the halogens as a Group are all quite reactive in their elemental forms, but their vigor declines below fluorine. Elemental iodine (I) is used as a common antiseptic and is sold in stores as "tincture of iodine"; the deep red-orange solution has elemental iodine dissolved in it. The most common compounds of the halogens are much tamer by comparison to the elemental forms, but some are not totally innocent. Fluorine compounds are used in toothpastes and in municipal water supplies to fight tooth decay. Teflon, used in pots, paints, plumbing and who knows what else, is a fluorine compound. Although these are mild examples, other fluorine compounds can eat glass or rock; these are sometimes used for etching ("frosting") designs into glass. Chlorine (Cl) compounds are very, very common. Normal everyday table salt is a chlorine compound. Many bleaching agents for use in home laundries and in home swimming pools are chlorine compounds and these are quite reactive.

Finally, Group 18. Right next to Group 17, the most reactive of all Groups, we have Group 18, the least reactive of all Groups. They're pretty much duds. Why? It's because of the way their electrons are arranged; we'll see this beginning in Chapter 22. Group 18 is called the "noble gases", although the terms "rare gases" and "inert gases" have also been used. They are not really rare, so that term has fallen out of favor. "Inert" is used in the normal dictionary sense and it very much reflects their chemical dudness, but other gases are also inert and that term has evolved into more general usage. The Group 18 elements are all gases and they form very few compounds. Helium (He) is the least reactive element of all. In fact, it is the least chemically reactive element or compound known in the universe. It forms no chemical compounds under anything close to normal conditions. Neon (Ne) is still pretty much lifeless. The Group does liven up as you go down further but you have to go to xenon (Xe) to start finding a reasonable number of compounds which stay around long enough to put in a bottle. Although this Group is inert, the elemental forms find various uses. Helium balloons float. Neon is used in the familiar red-orange neon signs. (Many of the other colors are fakes. They're not neon signs, although people call them neon signs. I'll explain this in Chapter 21.) If you take pictures and use a photoflash, you may be using xenon since it's in many flashtubes nowadays. At the bottom of the Group lies radon (Rn), which is chemically safe but is a radioactive hazard. Nature makes it continuously in the Earth but it doesn't stay there. It meanders through the ground, sometimes finding a crack in a home's foundation and ending up in a basement. There it constitutes a nuclear health hazard, and people who live in areas with high radon need to have their homes tested.

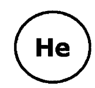
This concludes our Group tours.

Let me summarize at this point where we're at so you understand where we're going. This Chapter is about simple chemical units and their identities. Chemical identity begins with the atom. We built on that to give us the elements. The elements can exist by themselves in their elemental forms or they can form compounds with other elements. Compounds are different chemical units with different chemical identities. The simplest chemical units are composed of only one atom with nothing else joined to it. Those, however, are very few in number. Most chemical units have atoms joined to other atoms. This is true even for elemental forms, which I didn't tell you above. This is where we need to go right now.

2.6 Monatomic versus polyatomic: one versus many

When the chemical unit is composed of only one atom with no other atoms joined to it, we call this monatomic (shortened from mono-atomic). When the chemical unit is composed of two or more atoms joined together, we call this polyatomic. This category can be specified further using numerical prefixes such as diatomic (two atoms joined together), triatomic (three atoms joined together), etc. Just remember diatomic, triatomic, etc. are all versions of polyatomic. Monatomic is by itself.

There are very few substances on Earth for which the chemical unit is composed of only one atom at normal conditions. In fact, there are only six. They're the noble gases, Group 18, which I just mentioned above. They don't like to join with other atoms of any kind, including their own kind. Their elemental forms are monatomic. Thus, a helium balloon is full of individual helium atoms flying around inside, as represented by the simple picture at right.



All other substances on the planet exist as polyatomic chemical units composed of two or more atoms joined together. Except we don't say joined, we say bonded. We call the connection a "chemical bond". Chemical bonds are due to electrons. (Electrons do chemistry, remember?) There are different kinds of chemical bonds and we will talk a lot more on this in later Chapters. Chemical bonds can hold two or more atoms together in specific chemical units or they can extend over zillions of atoms in one, two or three dimensions to form a bonded "network" of atoms.

I can't tell you how important that last sentence is. Let me say it again.

CHEMICAL BONDS CAN HOLD TWO OR MORE ATOMS TOGETHER IN SPECIFIC CHEMICAL UNITS OR THEY CAN EXTEND OVER ZILLIONS OF ATOMS IN ONE, TWO OR THREE DIMENSIONS TO FORM A BONDED NETWORK OF ATOMS.

This one sentence sets the stage for every compound and every polyatomic elemental form in the universe. Don't forget it.

Now let's go to the next level for chemical unit and chemical identity: the molecule. A molecule is a neutral assembly of two or more atoms bonded together. The atoms can be of the same element or of different elements. Every molecule is identified by a specific number of specific atoms in a specific arrangement. This point is another biggie. Read it again:

EVERY MOLECULE IS IDENTIFIED BY A SPECIFIC NUMBER OF SPECIFIC ATOMS IN A SPECIFIC ARRANGEMENT.

Notice that there are three parts to a molecule's identity. "Specific number" is how many. "Specific atoms" refers to which elements. "Specific arrangement" includes shape such as plane, pyramid, diamond, etc. It also means which atoms are at what corners of the plane, pyramid, diamond, etc. ALL THREE PARTS TOGETHER DEFINE AND IDENTIFY THE CHEMICAL UNIT. All three together give rise to all the properties of that unit. IF YOU CHANGE ANY ONE OF THE THREE PARTS, THEN YOU CHANGE THE IDENTITY AND YOU HAVE A DIFFERENT CHEMICAL UNIT WITH A DIFFERENT SET OF PROPERTIES. This is extremely important.

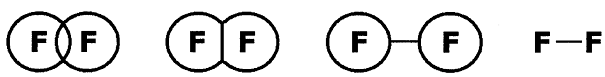
Water is water. A water molecule is composed of one atom of oxygen and two atoms of hydrogen. The shape involves the oxygen atom in the middle and the hydrogen atoms bonded to it at an angle. That's water's chemical unit. Every water molecule is that way. Every molecule of that type is identified as water.

But water is a compound, and I'm not at compounds yet. I need to step back just a bit and go back to elemental forms. (This moving around was explained in Chapter 1. Remember the Grand Puzzle. We're going for the big picture.)

We also need to bring in chemical formulas here.

When we were talking earlier about elements, I referred to elemental forms. We can now begin to identify specific cases of elemental forms. I already mentioned one such case when I said that Group 18 (and only Group 18) form monatomic chemical units: all the elemental forms for the noble gases are just the single atoms. Now let's do their formulas. In general, a chemical formula is used to identify a chemical (or other) unit using element symbols. For Group 18, the chemical unit is the single atom so the formula shows only the element's symbol. The elemental form for helium is designated by the formula He. The other five are just Ne, Ar, Kr, Xe and Rn. This may appear trivial, but that's the way it is for formulas of monatomic chemical units. Formulas for polyatomics can get ugly.

After monatomics, the next larger unit is the diatomic. All the halogens (Group 17) are diatomic in their elemental forms. Using fluorine as an example, we can represent this molecule pictorially in several different manners. The molecule is formed when the electron volumes overlap from the two atoms; this overlap is illustrated at far left. An equivalent illustration is provided by the second figure. The third figure uses a stick to represent the chemical bond. The far right figure is the simplest representation, again showing a stick for the chemical bond. All of these convey the same thing. All of these indicate that the molecule is composed of two atoms of fluorine joined by a chemical bond. Just remember that the left two pictures are the most realistic, since they convey the sense of overlap; stick drawings are often used, however, since their pictures tend to be clearer, especially when many atoms are involved.

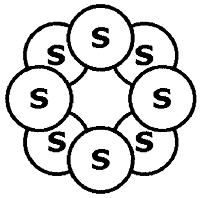


The chemical formula for this molecule is F_2. Subscripts are used in formulas to show the number of atoms of that element which are present if more than one. (Ones are not shown in a formula.) The chemical formulas for the other Group 17 elemental forms are Cl_2, Br_2, I_2 and At_2. In addition to the halogens, the common elemental forms for hydrogen, nitrogen and oxygen are also diatomic molecules. Their formulas are H_2, N_2 and O_2. I mentioned earlier that the atmosphere is mostly nitrogen and oxygen. We now see that these are molecules of N_2 and O_2.

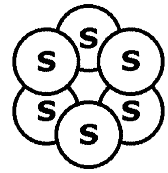
Out of the whole Periodic Table, only a few other elements have simple molecules as their elemental forms. Some of these elements have more than one form. For example, there's another elemental form for oxygen which is triatomic. This is ozone, O_3. It's a totally different chemical unit compared to O_2. It has a totally different set of properties. It's much more toxic than O_2 and will injure tissue and kill much faster. Nature makes it during lightning flashes and you can smell it after some thunderstorms. Nature also put a layer of it high in the atmosphere to protect surface critters from damaging ultraviolet radiation from the Sun. Unfortunately, O_3 near the surface of the Earth is a component of hazardous smog. So is ozone friend or foe? Like many things, it's both. But the key point for emphasis right now is how a change in the number of atoms of the chemical unit gives a very different molecule. O_3 is not O_2 and O_2 is not O_3.

By the way, O_2 and O_3 are "allotropes". Allotropes are different elemental forms of the same element. Allotropes have different chemical identities.

For other examples of molecular elemental forms, consider phosphorus and sulfur. One allotrope of phosphorus is P_4, which means there are four atoms of phosphorus bonded in the molecule. The four are in the shape of a pyramid. Sulfur forms a bunch of molecular allotropes. The most common is S_8, which means there are eight atoms of sulfur bonded in the molecule. The atoms form a zig-zag (up-and-



down) ring shape, as shown at left. There's also S_6 (shown at right), S_7, S_{12}, S_{18} and several more, most of which are ring shapes with different numbers of atoms in the ring. All of these are different molecules and they all have different identities.



So do we need to keep track of all of the various allotropes of many different elements? Sometimes a specific allotrope is indicated for a particular purpose and sometimes it is not. For example, we could specify S_8 for a particular purpose or we could simply say S which can then allow for different allotropes to be used. It just depends. Be aware that these things can happen.

So far, these are all examples of elemental forms which are simple molecules, but most of the elemental forms of most of the elements are "networks". I used that term earlier when I said "chemical bonds can hold two or more atoms together in specific chemical units or they can extend over zillions of atoms in one, two or three dimensions to form a bonded network of atoms". In a network, the bonds extend from one side of the sample to the other in one, two or three dimensions. Because of this, there's no simple molecule to point out. All metals form networks in their elemental forms. Some other elements also have network forms. (Networks are not limited to elements; many compounds such as quartz also do this.) In our stroll through the Periodic Table, I had mentioned that carbon is commonly graphite or diamond. What makes them different? They're both networks, but diamond has the bonds extending in three dimensions while graphite has them in only two dimensions. This makes a huge difference. (We'll talk a bit more of these differences in Chapter 38.)

Since there are no molecules within a network, we just use the element's symbol for the chemical formula. Thus, the elemental form for iron is just written Fe. The elemental form for gold is written Au. The elemental form for carbon is C, although this doesn't tell us which allotrope. This is OK for many uses, but if you do need to specify an allotrope you can include it in parentheses such as C(*diamond*).

Enough for elements. Now we're ready for compounds.

Problems

1. True or false.
   a. Most of the mass of the atom is in its nucleus.
   b. Electrons determine the size of an atom.
   c. The identity of an element is based only on the number of protons in its nucleus.
   d. All isotopes of one element have the same number of protons.
   e. Most elements have more than one isotope on Earth.
   f. Every atom of every isotope of Mo weighs 95.96 u.

2. True or false.
   a. Most elements are metals.
   b. Uranium is an inner transition element.
   c. Selenium is a metal element.
   d. Bromine is a halogen.
   e. Calcium is an alkali metal.
   f. Tin is in the Fifth Period.
   g. Neon is a noble gas.
   h. Aluminum is a Main Group metal element.
   i. Zinc is a transition metal element.

3. True or false.
   a. All halogens exist as monatomics as their normal elemental form.

b. O_2 and O_3 are allotropes of oxygen.
c. Every molecule is identified by a specific number of specific atoms in a specific arrangement.
d. Graphite is one allotrope of carbon.
e. In its typical elemental form, hydrogen is monatomic.

4. a. How many neutrons are in an atom of ^{196}Hg?
   b. How many protons are in an atom of ^{40}Ca?
   c. How many electrons are in an atom of ^{69}Ga?
   d. How many neutrons are in an atom of ^{32}S?
   e. How many protons are in an atom of ^{14}C?
   f. How many electrons are in an atom of ^{40}K?

5. a. What is the symbol of the halogen in the Fourth Period?
   b. What is the symbol of the noble gas in Period 3?
   c. What is the symbol of the Main Group metal with the smallest atomic number?
   d. In which Period is silver?
   e. In which Group is barium?
   f. What is the symbol of the alkali metal in the Fifth Period?
   g. What element is the least reactive element?
   h. What element is the most reactive element?

6. Silver comes in two isotopes naturally, ^{107}Ag and ^{109}Ag. ^{107}Ag has an abundance of 51.839% and a mass of 106.905092 u. ^{109}Ag has an abundance of 48.161% and a mass of 108.904757 u. Calculate the atomic mass of silver.

7. Gallium comes in two isotopes naturally, ^{69}Ga and ^{71}Ga. ^{69}Ga has a mass of 68.922580 u while ^{71}Ga has a mass of 70.924700 u. Calculate the percent abundance of ^{69}Ga. Use 69.723 u for the atomic mass of gallium.