"The periodic table."

CHAPTER 2. THE ELEMENTS: BUILDING BLOCKS OF THE UNIVERSE

Indestructible Elements, Indivisible Atoms

For centuries humankind has speculated about the nature of matter. So many different kinds of substances surround us. Are there simple principles that underlie this seeming complexity? Ordinary substances like water can be divided into smaller and smaller portions, tinier and tinier drops. Is there a point at which this is no longer possible? The Greek philosopher Democritus speculated in the fifth century B.C. that there must be a limit to divisibility. There must be tiny indivisible particles, or **atoms**, of a few basic substances, or **elements**, which combine in different proportions to make up the substances of the universe. According to the Greek idea, there were four basic elements: fire, air, earth, and water. Not all Greek philosophers were in agreement about these ideas; the influential philosopher Aristotle accepted the elements, but rejected the concept of the atom on mathematical grounds.

In the first century B.C. Lucretius, a Roman philosopher of the Epicurean school, included atomism as an important part of his lengthy poetic work *De rerum natura*, in which he declared that all matter in the universe, even humankind itself, is made of immutable and indivisible atoms. Lucretius was important, not only as a philosopher, but also as an early science writer. Throughout history those individuals who can both understand science concepts and explain them to the public at large have been in short supply. Lucretius succeeded in explaining the concept of the atom through poetry:

The primal elements must needs be made
Of stuff immortal, whereinto all things
At their last hour can be dissolved, that so
Their matter for the birth of other things
May be supplied. These primal bodies, then,
Are single, solid, indivisible:
Nor could they else throughout the eternal roll
Of endless time now gone have been preserved,
And of their substance build all things anew...
There are, then,
These unseen bodies, strong in singleness,
Whose union, when close-packed, can cause all things
To stand firm knit and show their stalwart strength.

De rerum natura is a lengthy work which includes not only the Epicurean concept of atomism, but a defense of science over superstition as well. If substances are formed by the harmonious combination of atoms acting according to natural laws, and indeed if all nature behaves according to natural laws which can be scientifically studied, what role is left to be played by the capricious and supposedly omnipotent gods? The ideas of Lucretius were considered to be in opposition to the established Roman religion of his day; early Christianity thought his views heretical as well. Not until scientific inquiry began to flourish in the Renaissance was the importance of atomic theory recognized.

Because the atom is so very small, even twentieth-century scientists had to rely on logical arguments as Lucretius did. Although the concept of the atom was universally accepted, no one had actually seen an individual atom until almost the end of the twentieth century. In 1981 Gerd Binnig and Heinrich Rohrer at the IBM Research Laboratory in Zurich, Switzerland, invented the scanning microscope, with which they were not only able to see individual atoms, but move them around. In the April 5, 1990, issue of the scientific journal *Nature* readers were treated to an astounding and historic photo: researchers at the IBM Almaden Research Center in California had been able to move around individual atoms of xenon on a smooth nickel surface with such precision that 35 xenon atoms were arranged in a crisp and unmistakable pattern: the IBM logo.

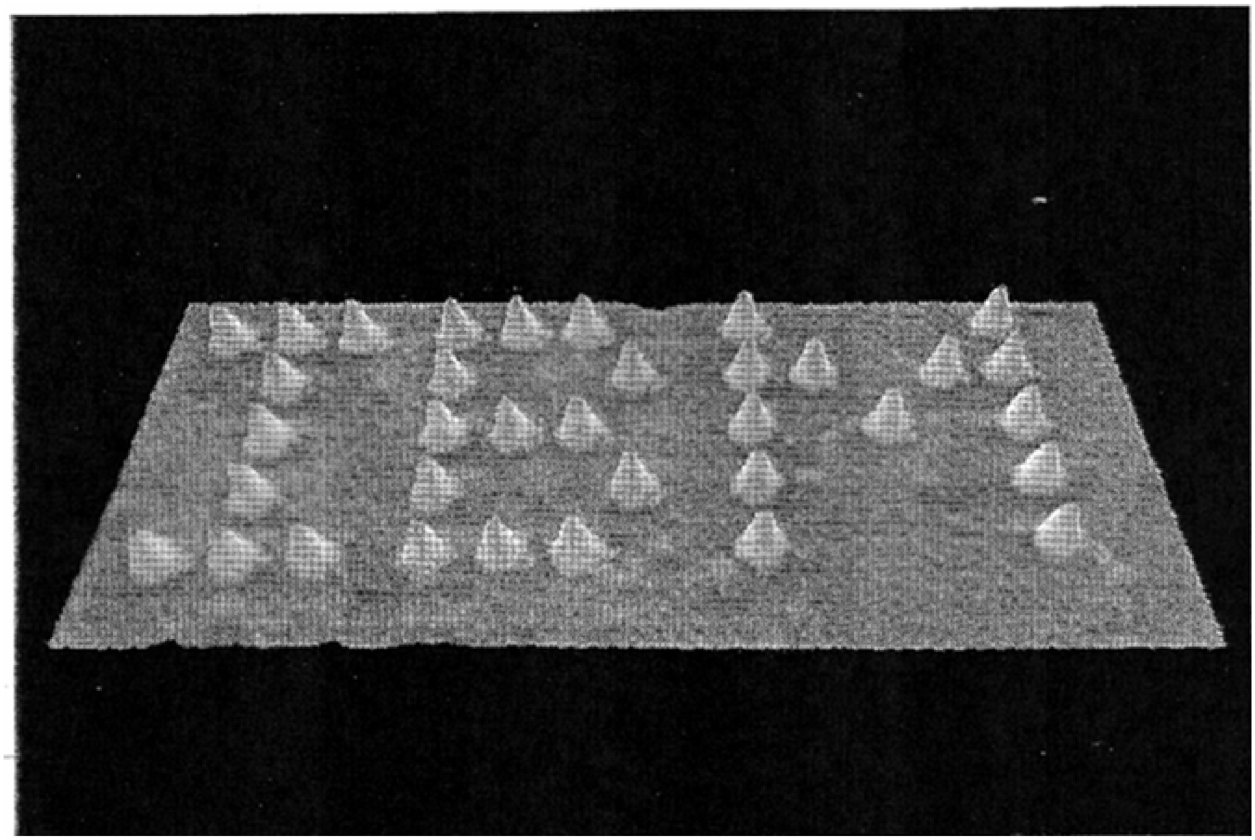


Fig. 2-1. In 1990 the scanning tunneling microscope produced this photo of 35 individual xenon atoms arranged to form the IBM logo.

The Greeks and Romans were philosophers, not scientists. They reasoned from their observations of nature, but did not conduct laboratory experiments in which unproven suppositions, or hypotheses, were tested by careful measurements. The early scientists who studied matter soon began to reject the idea that the elements were the four which had been proposed by the Greeks: fire, air, earth, and water. The Englishman Robert Boyle, in his book *The Skeptical Chymist*, published in 1661, suggested that all substances could be tested to determine whether they were elements. True elements could not be broken down into simpler substances no matter how the chemist tried. If a substance could be broken down into simpler substances, then it was a **compound** of those elements.

In 1789 the relatively new science of chemistry advanced enormously with the publication of what has been called the first chemistry textbook. The French scientist Lavoisier summarized the results of his experiments along with those of others in the highly influential *Traite elementaire de chimie*, or *Elementary Treatise on Chemistry*. His clear and compelling logic vanquished several erroneous theories and established firmly the concept of elemental substances as determined by experiment. He listed in a table those substances he believed to be elements, the majority of which appear in our table of elements today.

Chemical Building Blocks: a Limited Set

The simple yet powerful concept of the elements, each composed of characteristic atoms, is the basis of much of the science of chemistry. The number of elemental substances which make up the earth is limited, and the ways in which elements react together to form compounds follows logical patterns. The numerous chemical substances which form the matter of our everyday lives, the household products, drugs, pollution, even our own bodies, are all constructed from this very limited set of building blocks, the elements. Even when we travel outside our planet, visiting the moon or sending probes to outer space, our chemical analysis need seek for only those familiar elements in our limited list.

For the nonscientist, it should be comforting to know that the bewildering array of chemicals and chemical issues that confront us today is part of an organized system with quite a manageable number of parts. Understanding some basic chemical principles can lead to understanding of a wide variety of issues. Mastery of these principles is possible for the average person, and is a welcome alternative to the feelings of inadequacy and helplessness that are sometimes caused by the seemingly complex issues of today's technological society.

Just how long is the list of all the elements? Table 2-1 lists all the elements. You will notice the total number of elements in the list is 109, not bad considering that all the substances in the universe can be assembled out of combinations from this list. The true list of "usable" elements is even shorter. The elements at the very end of the list have been produced with great difficulty by scientists in recent years; they are unstable, with lifetimes sometimes as short as seconds. The official names of these last few elements on the period table have not yet been set, and are currently a subject of debate in international scientific societies.It has been a great challenge to characterize these ephemeral elements, but we hardly need to concern ourselves with these elements as basic building blocks of our everyday lives. The transuranium elements, or those higher in atomic number than uranium, number 92, are unstable, radioactive elements that will be mentioned later in connection with radioactivity.

Table 2-1: The elements, their symbols, and their atomic numbers

1	Hydrogen	H	42	Molybdenum	Mb	83	Bismuth	Bi
2	Helium	He	43	Technetium	Tc	84	Polonium	Po
3	LithiumLi		44	Ruthenium	Ru	85	Astatine	At
4	Beryllium	Be	45	Rhodium	Rh	86	Radon	Rn
5	Boron	B	46	Palladium	Pd	87	Francium	Fr
6	Carbon	C	47	Silver	Ag	88	RadiumRa	
7	Nitrogen	N	48	Cadmium	Cd	89	Actinium	Ac
8	OxygenO		49	Indium	In	90	Thorium	Th
9	Fluorine	F	50	Tin	Sn	91	Protactinium	Pa
10	Neon	Ne	51	Antimony	Sb	92	Uranium	U
11	SodiumNa		52	Tellurium	Te	93	Neptunium	Np
12	Magnesium	Mg	53	Iodine	I	94	Plutonium	Pu
13	Aluminum	Al	54	Xenon	Xe	95	Americium	Am
14	Silicon	Si	55	Cesium	Cs	96	Curium	Cm
15	Phosphorus	P	56	Barium	Ba	97	Berkelium	Bk
16	Sulfur	S	57	Lanthanum	La	98	Californium	Cf
17	Chlorine	Cl	58	Cerium	Ce	99	Einsteinium	Es
18	Argon	Ar	59	Praseodymium	Pr	100	Fermium	Fm
19	Potassium	K	60	Neodymium	Nd	101	Mendelevium	Md
20	Calcium	Ca	61	Promethium	Pm	102	Nobelium	No
21	Scandium	Sc	62	Samarium	Sm	103	Lawrencium	Lr
22	Titanium	Ti	63	Europium	Eu	104	Kurchatovium	Ku
23	Vanadium	V	64	Gadolinium	Gd	105	Hahnium	Ha
24	Chromium	Cr	65	Terbium	Tm	106	Seaborgium	Sg
25	Manganese	Mn	66	Dysprosium	Dy	107	Bohrium	Bh
26	Iron	Fe	67	Holmium	Ho	108	Hassium	Hs
27	Cobalt	Co	68	Erbium	Er	109	Meitnerium	Mt
28	Nickel	Ni	69	Thulium	Tm			
29	Copper	Cu	70	Ytterbium	Yb			
30	Zinc	Zn	71	Lutetium	Lu			
31	Gallium	Ga	72	Hafnium	Hf			
32	Germanium	Ge	73	Tantalum	Ta			
33	Arsenic	As	74	Tungsten	W			
34	Selenium	Se	75	Rhenium	Re			
35	Bromine	Br	76	Osmium	Os			
36	Krypton	Kr	77	Iridium	Ir			
37	Rubidium	Rb	78	Platinum	Pt			
38	Strontium	Sr	79	Gold	Au			
39	Yttrium	Y	80	Mercury	Hg			
40	Zirconium	Zr	81	Thallium	Tl			
41	Niobium	Nb	82	Lead	Pb			

.

If we were to analyze the elements in the substances we encounter every day, we would find that the basic list is more limited still. Just like a child's set of building blocks, which can be purchased with a few fancy pieces that will make a chimney top or a window frame or shutters, but still relies on the repetitive use of a few basic pieces to build the walls of a structure, the atomic building set uses a few elements over and over to make most structures. Our bodies, for example, are made mostly of the elements carbon, hydrogen, oxygen, and nitrogen. (See Table 2-2, Elements Found in the Human Body.) The earth's crust is made mostly of the elements oxygen and silicon. (See Table 2-3, Elements Found in the Earth's Crust.) Over 99% of the earth's atmosphere is made of the elements nitrogen and oxygen. The major components of the Earth's atmosphere are listed in Table 2-4.

Table 2-2. Elements Found in the Human Body*

Element	*Abundance*
Hydrogen	63%
Oxygen	25.5%
Carbon	9.5%
Nitrogen	1.4%

**Does not include trace elements!*

Table 2-3: Elements Found in the Earth's Crust

Element	*Abundance*
Oxygen	60.1%
Silicon	20.1%
Aluminum	6.1%
Hydrogen	2.9%
Calcium	2.6%
Magnesium	2.4%
Iron	2.2%
Sodium	2.1%

The First Twenty Elements

The first twenty elements include all those that are listed as the predominant elements in the earth's crust and atmosphere and in the human body, as well as many of the most important elements from a chemists's point of view. We will learn that other elements, the transition metals like iron and copper, for example, have important roles. Most chemical concepts we will learn, however, can be illustrated from this very short list of twenty elements, shown in Table 2-4.

Table 2-4. The First Twenty Elements

Atomic Number	*Name*	*Symbol*	*Gram Atomic Mass*
1	Hydrogen	H	1.008
2	Helium	He	4.00
3	Lithium	Li	6.94
4	Beryllium	Be	9.01
5	Boron	B	10.8
6	Carbon	C	12.0
7	Nitrogen	N	14.0
8	Oxygen	O	16.0
9	Fluorine	F	19.0
10	Neon	Ne	20.2
11	Sodium	Na	23.0
12	Magnesium	Mg	24.3
13	Aluminum	Al	27.0
14	Silicon	Si	28.1
15	Phosphorus	P	31.0
16	Sulfur	S	32.1
17	Chlorine	Cl	35.5
18	Argon	Ar	39.9
19	Potassium	K	39.1
20	Calcium	Ca	40.1

Each of the elements has a symbol. These symbols will be convenient shorthand as we begin to discuss chemical structures. Notice in Table 2-4 that the first letter of the symbol is always capitalized, but the second letter is not. Most of the symbols are logical, H for hydrogen, for example, and He for helium. Some other symbols seem less obvious, like Na for sodium. Pb for lead and Au for gold are other such examples from the full list of the elements. Interesting stories lie behind these unlikely symbols. Long before the modern science of chemistry developed in the eighteenth century, chemical substances were being used and investigated for their useful properties. The Egyptians used sodium compounds and other chemical substances as part of their preservative formulas in mummification. The Greeks., Arabs, and finally the Medieval Europeans inherited this tradition of working with chemical substances, which became known as alchemy. The alchemists are most often remembered for their search to find the "philosopher's stone" which would change lead into gold. Hence they are not considered true scientists, for they did not use the scientific method. Instead of using experiments to prove or disprove the hypothesis that lead can be changed into gold, they began with the assumption that it could be done, and spent their lives trying to do so. For them, a "good" experimental result was not finding out the truth, but getting the results they wanted. However, many of the techniques they developed to purify and identify substances were useful to the scientists who followed them and generated the theories of chemistry. The symbols for lead, gold, and sodium predate modern chemistry and derive from their old Latin names, plumbum, aurum, and natrium. Thus, as we write the chemical formulas for these elements we acknowledge the role of these ancient roots of chemistry.

Why are the elements arranged in this particular numbered order? Finding the best way to organize the list of elements has been one of the major problems of chemistry since its early days. As we learn more about the elements, the answer to this question will become clearer, and we will return to it as we gain more chemical information.

Elements and Compounds, Atoms and Molecules

The concept of the atom was generally accepted by scientists much later than that of the elements. After all, these smallest possible particles are much too small to be seen by anyone; their existence was far from obvious. The atomic theory was accepted only after decades of experimental work by many scientists observing how the elements combined by weight. It is an excellent example of what scientists mean by the word *theory*. An inspired guess, or hypothesis, about the laws of the natural world, must be tested thoroughly by experiment before it is accepted and finally attains the status of a theory.

Sometimes the elements occur naturally in pure form, uncombined with other elements. They are much more likely to be found combined with other elements in substances called **compounds**. As the Greek and Roman philosophers first explained, it is the combination of the elements in many different types of compounds that makes possible the infinite variety of substances in our universe.

To be able to discuss how all the many compounds in the world can be formed from a relatively few elements, let us first define some terms:

> An **atom** is the smallest possible unit of an **element**.
>
> **Elements** combine to produce **compounds**.
>
> The smallest possible unit of a **compound** is called a **molecule**.

A molecule is made up of atoms of its constituent elements. For example, a molecule of hydrogen chloride, represented as HCl, is composed of a hydrogen atom and a chlorine atom. A molecule of water, represented as H_2O, is composed of two hydrogen atoms and an oxygen atom.

Fig. 2-2. *The hydrogen chloride molecule and the water molecule*.

The hydrogen and oxygen in water always appear in this 2:1 ratio; this constant ratio is one of the proofs by which early chemists were able to recognize chemical compounds and distinguish them from mixtures of substances. (Of course, the properties of water are very different from the two gases hydrogen and oxygen, another indication that it is more than simply a mixture of hydrogen and oxygen.) In contrast, for example, salt water is not a compound, but a mixture of salt (sodium chloride) and water, which can be mixed in any proportions.

DECODING CHEMICAL FORMULAS

Chemical formulas use the element symbols to show which elements are present in a compound. They also tell how many atoms of each element are found in one molecule of the compound. If more than one atom of an element is present in a molecule, a subscript to its right tells how many atoms. If no subscript is present, assume that one atom of that element is present. For example, the formula for water is H_2O, indicating that the water molecule is made up of two hydrogen atoms and one oxygen atom.

Two hydrogen atoms → H_2O ← One oxygen atom

Table 2-4. Composition of the Earth's Atmosphere at Sea Level

Gas	***Percentage by Volume***
Nitrogen (N_2)	78.084
Oxygen (O_2)	20.948
Argon (Ar)	0.934
Carbon dioxide (CO_2)	0.033
Neon (Ne)	0.00182
Helium (He)	0.00052
Methane (CH_4)	0.0002

Problem example 2.1.
DECODING SKILLS: ELEMENTS AND COMPOUNDS IN THE EARTH'S ATMOSPHERE

In the list of the major components found in the earth's atmosphere given in Table 2-4, Which are elements? Which are compounds?

From the chemical formulas we see that nitrogen (N_2), oxygen (O_2), argon (Ar), neon (Ne), and helium (He) are not combined with other elements, and hence they occur in the atmosphere as pure elements, not compounds. Notice that argon, neon, and helium occur as single atoms, while from their chemical formulas it is apparent that nitrogen and oxygen each have two atoms joined together; they are sometimes referred to as "***diatomic*** gases."

Carbon dioxide (CO_2) is a compound; its formula shows that a carbon dioxide molecule is made from one carbon atom and two oxygen atoms. Methane (CH_4) is also a compound; its chemical formula shows that the methane molecule is made from one carbon atom and four hydrogen atoms.

Decoding Skills in the Supermarket; Using the Element Concept

At this point we have not learned a great deal about chemistry, but already we have some useful information to apply to everyday life. For instance, since sodium appears on the list of the elements, we can expect there will be a great number of sodium compounds, all made from this basic building block. Later we will learn about the type of compounds characteristic to sodium; they are called ionic compounds, or salts. When for health reasons, perhaps related to hypertension or fluid retention in the body, people are advised to cut down on salt consumption, it is sodium compounds that are actually being referred to. Not only table salt, or sodium chloride, but any other sodium compounds such as sodium citrate or monosodium glutamate are sources of sodium.

Fig. 2-3. Photo of label of Kraft macaroni and cheese box.

Problem Example 1-1: In the following list of ingredients found on a box of macaroni and cheese, identify all the compounds that are listed as sources of sodium:

> *Ingredients: enriched macaroni, cheese sauce mix [whey, dehydrated cheese (granular and cheddar milk cheese culture, salt, enzymes)], whey protein concentrate, skim milk, salt, buttermilk, sodium tripolyphosphate, sodium phosphate, citric acid, yellow 5 & 6 (color), lactic acid.*
>
> Sodium chloride is listed twice as "salt", once in the cheese sauce mix, and again as a separate ingredient. Sodium tripolyphosphate and sodium phosphate are sodium compounds also.

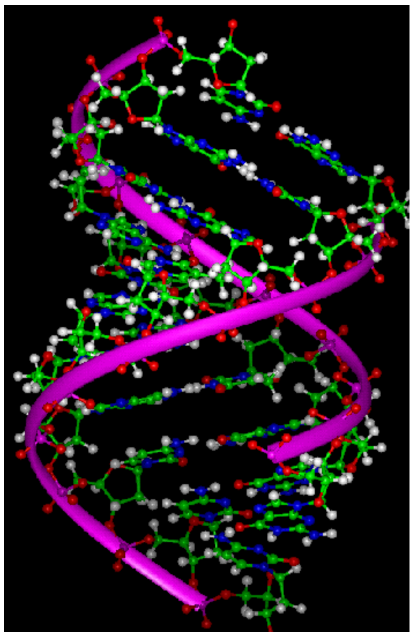
Solving this simple problem may have introduced some additional ideas beyond those required to identify compounds of sodium. It may have occurred to you, for instance, that this seems to be rather a large amount of sodium. Nutritional information on the box states that a 3/4 cup serving of the prepared macaroni and cheese dinner contains 530 mg of sodium. This, however, is the weight of sodium alone. In later chapters it will be possible to calculate the weight of table salt, or sodium chloride, that would correspond to this amount of sodium. For 1.06 grams of sodium, it is 2.7 grams! You might want to weigh out this amount on a laboratory balance if you are not yet familiar enough with metric weights. (If you are more likely to eat 1 1/2 cups of macaroni and cheese, try weighing out 5.4 grams.)It is the author's hope that it may also have entered your thoughts that it may be useful to learn yet a bit more chemistry, since we have only begun to decipher the many mysteries of the macaroni and cheese box.

How Can We Look at a Molecule? The Role of the Molecular Model

One reason why the behavior of atoms can be difficult to comprehend is the fact that atoms and molecules are so very small; they can never be seen with the naked eye. If the water molecule is difficult to visualize, how can we hope to have a clear picture of the much more complex molecules that make up much of our own bodies and our world? Even the most learned scientists confront this problem. In discussing spatial visualization skills in Chapter 1 we have shown how chemists use **molecular models** mentioned the tools used to solve this problem.

The molecular models used by chemists to represent the shapes of atoms and molecules can vary widely. They can be simple and inexpensive, for example, constructions made of gumdrops or styrofoam balls, or complex and expensive, like computer-generated images of DNA molecules that can rotate to show three-dimensional character. This text, like almost all books that include descriptions of molecules, will sometimes use pictures of molecular models. Some types of models are useful in showing the shapes of molecules. Others show the bonds holding the molecules together. All pictures of models in this book, however, be as flat and two-dimensional as they page on which they sit, and will lack the three-dimensional quality which is one of the major advantages of molecular models.

Fig. 2-4. *Molecular models can be simple or complex. At left is a simple molecular model set that can be purchased at Target; at right is a model of DNA.*



Atomic Mass

Water, and its component elements hydrogen and oxygen, figured prominently in the experiments that established the atomic theory. The chemistry of the compound water, or H_2O, still holds an important role in our lives, and frequently we will use it to illustrate an important chemical principle. Here, the concept of **atomic mass**, central to chemical science, is illustrated by the combination of hydrogen and oxygen to form water.

Hydrogen and oxygen are both gases; they react together violently in the presence of a spark to produce water. Chemists who perform this experiment observe that invariably they react in the same proportions by weight. Hydrogen reacts with about eight times its weight in oxygen. The English chemist John Dalton in the early nineteenth century had the insight to realize the monumental implications of this simple fact . The only simple explanation for this recurring whole-number weight relationship in a chemical reaction, which is observed in other reactions as well, is that the ancient theory of atomism is correct: elements are in fact made up of atoms; moreover, **different atoms have different characteristic masses**. Dalton assumed that the formula of water is HO, and that therefore the relative weights, or, more properly, the relative masses, of hydrogen and oxygen are one and eight. (Actually, the value he found for oxygen was closer to 7.5.) Later, other chemists were able to deduce from a series of experiments that the formula for water is H_2O, with two hydrogen atoms for every oxygen atom. The relative weights of hydrogen and oxygen must be one and sixteen. In other words, **an oxygen atom is sixteen times as heavy as a hydrogen atom.**

Fig. 2-5. *The oxygen atom is sixteen times as heavy as the hydrogen atom.*
(Possible illustration: the two spherical atoms, labelled with their masses. Possibly, on a double-pan balance with the oxygen sitting lower showing it's heavier.)

This important principle was a major breakthrough in the study of chemistry. Not only was the theory of atomism validated, but a very important individual property of the atoms of a given element, their relative mass, was discovered. John Dalton found a way to assign a mass value to each element ; he simply assigned an arbitrary value of one to the lightest element he had found, which was hydrogen. Oxygen, then, which is sixteen times as heavy as hydrogen, is given a mass value of sixteen. By this time scientists were using the gram as their unit of mass measure, as we do today. So they found it convenient to express atomic mass units in terms of **gram atomic mass**, with the gram atomic mass of hydrogen being one gram and the gram atomic mass of oxygen being sixteen grams. Fortunately, hydrogen, the lightest element Dalton knew of, is actually the lightest element of all. So the atomic masses discovered in the earliest days of modern atomic theory are very close to the ones in use today.

On the periodic table (inside cover of the text) the gram atomic mass of each element is shown, as well as the atomic numbers. Notice that the gram atomic masses are always larger than the atomic numbers. Unlike the atomic numbers, the gram atomic masses are not whole numbers. Chapter 3, *Inside the Atom*, explains why this is so. Chapter 5, *Chemical Reactivity and the Periodic Table*, explains the arrangement of the rows and columns of the periodic table.

Look again at the first twenty elements (either in the periodic table or in Table 2-4, comparing the atomic masses. The order in the list is almost identical to the order by increasing mass. Why isn't the order perfect? Not until we enter the twentieth century of scientific discovery in Chapter 3 will we know.

How Small Is an Atom? Avogadro's Number

We already know that an atom is far too small to be seen with the naked eye, but just how small is an atom? Scientific experiments show that the number of atoms in one gram of hydrogen is 6.02×10^{23}. An oxygen atom, you will recall, is sixteen times heavier than a hydrogen atom. Therefore sixteen grams of oxygen has the same number of atoms as one gram of hydrogen, 6.02×10^{23} atoms. Applying this concept in a more general way, we can say that **one gram atomic mass of any element contains the same number of atoms, 6.02×10^{23} atoms**. This fact is an important aspect of atomic theory. It means that we can conduct chemical reactions on a gram scale yet still react equal numbers of atoms with each other. The number 6.02×10^{23} is given the name **Avogadro's number** in honor of the Italian scientist Amadeo Avogadro, a contemporary of Dalton who made major contributions to the theory of atomic mass.

Although an important component of chemical theory, Avogadro's number is not a number used daily by chemists like the measurement units of grams and milligrams; in fact, its actual value was not determined until after gram atomic masses were in common use. For nonscientists, understanding the magnitude of Avogadro's number can be useful as a key to appreciating just how small an atom is. How can we comprehend a number of this size?

$$6.02 \times 10^{23} = 602{,}000{,}000{,}000{,}000{,}000{,}000{,}000$$

If you were able to spend a million dollars a second, it would take twenty billion years to spend Avogadro's number of dollars! Or, if you were able to count a million atoms a second, it would take twenty billion years to count to Avogadro's number, the number of atoms in one gram atomic weight.

Checking Your Understanding of the Concept of Avogadro's Number Avogadro's number (6.02×10^{23}) is a fact for you to memorize.

If you understand the concept of Avogadro's number, you should be able to fillin the blanks easily, without doing calculations!

One gram atomic mass of any element contains ___________ atoms.

4.0 grams of helium contains ___________ atoms.

16.0 grams of oxygen contains ___________ atoms.

8.0 grams of oxygen contains ___________ atoms.

Problem example 2-2: How many atoms are present in 28.0 g of nitrogen?

In order to solve this problem, we will use the fact that 6.02×10^{23} atoms are present in one gram atomic mass of any element. The gram atomic mass of an element can be found on any periodic table of the elements, and the gram atomic masses of the first 20 elements are listed in Table 2-4. First, let us reason through this problem stepwise:

A gram atomic mass of nitrogen has 6.02×10^{23} atoms.

A gram atomic mass of nitrogen has 14.0 grams.

Therefore, 14.0 grams of nitrogen have 6.02×10^{23} atoms.

28.0 grams of nitrogen have twice as many atoms as 14.0 grams of nitrogen, or $2 \times 6.02 \times 10^{23}$ atoms = 12.04×10^{23} atoms.

Or, analyzing the problem by the conversion factor method, and setting up the samew reasoning in a mathematical format,
we have the number of grams, but **we want** the number of atoms.

We have	*We want*
28.0 g nitrogen	atoms of nitrogen

Setting up the conversion factor, we see that atoms of nitrogen will be on the top of the conversion factor, and grams of nitrogen will be on the bottom, where it will cancel out.

$$28.0 \text{ g nitrogen} \times \frac{\text{atoms of nitrogen}}{\text{g of nitrogen}} = \text{atoms of nitrogen}$$

A conversion factor which relates the number of grams to the number of atoms would make possible the solution to the problem. Avogadro's number, 6.02×10^{23} atoms in 1 gram atomic mass, combined with the number of grams in 1 gram atomic mass, provides us with the conversion factor. This can be done in two steps or in one step as shown below:

$$28.0 \text{ g nitrogen} \times \frac{1 \text{ gram atomic mass}}{14.0 \text{ g nitrogen}} \times \frac{6.02 \times 10^{23} \text{ atoms of nitrogen}}{1 \text{ gram atomic mass}} = \text{atoms of nitrogen}$$

$$28.0 \text{ g nitrogen} \times \frac{6.02 \times 10^{23} \text{ atoms of nitrogen}}{14.0 \text{ g of nitrogen}} = \text{atoms of nitrogen}$$

Whether you choose to set up this problem in one step or two steps, the problem-solving technique is the same as outlined in Chapter 1:

1. Write down what information is given in the problem ("*we have*") on the left side of the page, and what the goal of the problem is ("*we want*") on the right side. Be careful to include all unit labels, as these are important components of the given information and the problem goal.

2. Set up the conversion factor that will give the correct answer by writing on the top of the factor the units for what you want, and on the bottom (where they will cancel out) the units for what you were given. Use your chemical knowledge to supply the correct numbers for the conversion factor. For example, the number that relates the number of atoms to the mass is Avogadro's number, 6.02×10^{23} atoms in one gram atomic mass.

3. Check to see whether all units cancel out, giving the desired units for the answer. If not, can you add another conversion factor, solving the problem in two operations.

4. Finally, solve the arithmetic and check to see whether the answer makes sense. In this case, the answer is 12.0×10^{23} atoms of nitrogen. This is not only about the magnitude of number we would expect for such a problem, but makes sense since we have twice the mass in one gram atomic mass, and expect twice Avogadro's number of atoms.

Problem example 2-3: How many grams does one atom of nitrogen weigh?

In this problem, *we have* 1 atom of nitrogen, and *we want* to know how many grams.
Writing down this information.

we have	*we want*
1 atom of nitrogen	grams

Setting up the conversion factor, we see that grams will be on top of the conversion factor, and atoms of nitrogen on the bottom.

$$1 \text{ atom of nitrogen} \quad \times \quad \frac{\text{grams}}{\text{atoms of nitrogen}} \quad = \quad \text{grams}$$

As in the some of the metric conversion problems in Chapter 1, this problem could be solved in two steps. We may not know the number of grams in an atom, but we do know the number of grams in one gram atomic mass. We also know the number of atoms in one gram atomic mass. Putting in what we do know, we find that units cancel out to give what we want to know.

$$1 \text{ atom of nitrogen} \quad \times \frac{1 \text{ gram atomic mass}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{14.0 \text{grams of nitrogen}}{1 \text{ gram atomic mass}} = \text{grams of}$$

Solving the arithmetic, we find the answer, 2.33×10^{-23} grams.

Checking to see whether the answer makes sense, we see that this answer is very small, and we would expect the mass of an atom to very a very small number.

For those who are familiar with the concept of Avogadro's number (see "Checking your understanding of the concept of Avogadro's number"), it may be simpler to solve in one step, since 6.02×10^{23} atoms of nitrogen weigh 14.0 grams. Hence,

$$1 \text{ atom of nitrogen} \quad \times \frac{14.0 \text{grams of nitrogen}}{6.02 \times 10^{23} \text{ atoms}} = \text{grams}$$

Electromagnetic Radiation: Not Matter, But Energy

The atomic theory gives us a basis for understanding the basic structure of all the substances, or matter, in the universe. But there are phenomena in the universe that cannot be described as matter. Matter has mass and occupies space. Its mass is subject to gravitational attraction, which we utilize in weighing it. Objects in space may float free because they are weightless, that is, free of gravitational attraction. Yet they still have mass and occupy space. Some natural phenomena, however, cannot be described as matter. Movement, heat, light, and radio waves are examples of **energy**; we can experience their existence, but we cannot weigh them as we can weigh matter, for they have no mass we can measure, nor do they occupy space as matter does.

The dawn of the twentieth century brought some spectacular new insights into the nature of atoms. These discoveries were made possible in part by experiments which utilized the type of energy called **electromagnetic radiation**.

Perhaps the form of electromagnetic radiation we are most familiar with is **visible light**. It illuminates our days with brightness and color, and when night falls we use man-made sources of the visible light we have come to take for granted. Just what is this commonplace phenomenon on which we depend so much? Experimenting with the object called a prism can help us to understand the nature of visible light. The prism can be used as a tool or a toy; when it is held at the correct angle to a beam of visible light, it transforms the white light into a rainbow of colors. Droplets of water in the sky under the right conditions can do the same thing, giving rise to the rainbow in the sky. The colors of the rainbow always appear in the same order, shading from violet through blue and green to yellow, then to orange and last of all to red. Each color corresponds to an energy value; violet is the highest-energy light of the visible spectrum, and red is the lowest.

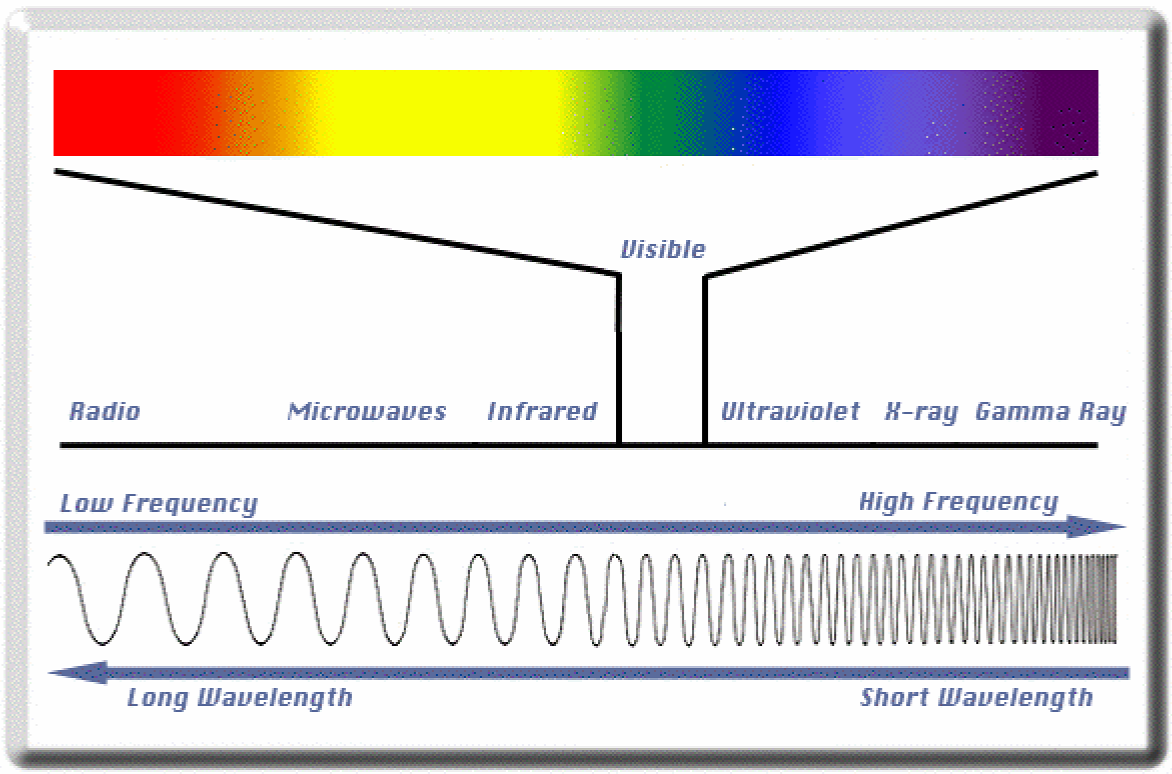


Fig. 2.6. *The electromagnetic spectrum.*

The visible light which we can see, however, does not comprise all of the electromagnetic spectrum. Beyond the violet is light of even higher energy, the **ultraviolet**. Though human eyes cannot respond to ultraviolet radiation, insects have a different range of perception. The bee, for instance, cannot see red, but can see yellow and blue shades, and can distinguish color into the ultraviolet range. As is evident from Figure 2-7, the bee's eye view of the marsh marigold is very different from our own. It is light within the ultraviolet range which is high enough in energy to give us a sunburn. That is why it is possible to get a sunburn on a cloudy day; the ultraviolet radiation may be penetrating the clouds, even though the maximum amount of visible light is not shining through. Like visible light, ultraviolet light exists as a continuum of energies. The near ultraviolet, or UVA, wavelength range, is lowest in energy; UVB or mid-UV and UVC or far UV are higher in energy and more harmful.

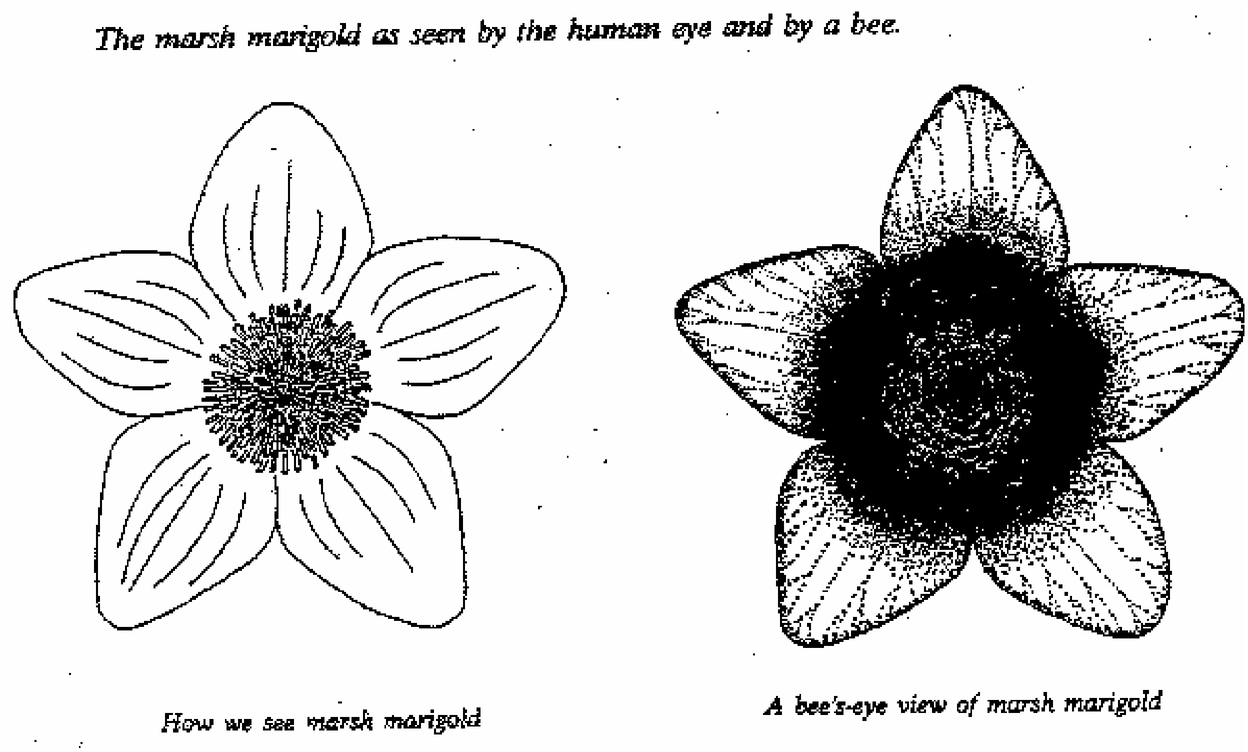


Fig. 2.7. *The marsh marigold as seen by the human eye and by a bee. (from Mass Audubon Society)*

Beyond the red end of the visible spectrum lies the range of electromagnetic radiation we call the **infrared**. Infrared energy cannot be seen by the human eye, but we feel this form of energy as heat. An infrared heat lamp glows in the red-to-infrared region; we can see the red and feel the infrared energy given off by the lamp. Infrared cameras have numerous practical uses. An infrared photo of a house on a cold day will show as "hot spots," or dark areas on the photo, those places in the exterior of the house where heat is being lost, for instance, the windows or areas with insufficient insulation. Such photos can be useful in making homes more energy-efficient. Infrared cameras have important military applications in wartime, as hot engines and even human bodies are plainly visible as infrared radiators even at night.

There are other forms of electromagnetic radiation as well. X-rays and gamma rays, which will soon play a part in our discussion of early experiments to investigate atomic structure, are very high-energy forms of radiation. Microwaves are lower in energy than infrared, and radiowaves are lower in energy still. All these phenomena, which may at first seem unrelated, are all simply forms of electromagnetic radiation.

Associated with each energy value for electromagnetic radiation are values of **wavelength** and **frequency**. The simple equations that relate these properties show that there is a direct relationship between energy and frequency; that is, the higher the energy value, the higher the frequency will be. There is an inverse relationship between energy and wavelength; that is, the higher the energy value, the shorter the wavelength will be. In equation form,

$$E = h\nu \qquad E = \frac{hc}{\lambda}$$

$$E = energy$$

$$\nu = \ frequency$$

$$\lambda = wavelength$$

$$h = P'lancks\ constant$$

$$= 0.6626\ x\ 10^{-}33\ \ joule - \mathrm{sec}$$

$$c = speedoflight$$

$$= 3.00\ x\ 10^{8}\ m$$

Problem example 2-4: Since radio waves are lower in energy than visible light, which has a longer wavelength, visible light or radio waves?

Since there is an inverse relationship between energy and wavelength (higher energy, shorter wavelength), the radio waves have a longer wavelength. To give an idea of their relative magnitudes, a typical wavelength for visible light is about 5×10^{-7} meters; a typical radio wave could be 3×10^{3} meters.

Problem example 2-5: Which has a higher frequency, visible light or radio waves?

Fig. 2.5 shows that visible light is higher in energy than radio waves. There is a direct relationship between energy and frequency; the higher-energy visible light also has a higher frequency.

To get a more concrete example of these theoretical concepts of energy, frequency, and wavelength, you can think in terms of the standing wave that can be generated by moving a rope like a child's jump rope in a vertical plane, keeping the ends in the same position at all times. Simply moving the center of the rope up and down creates one half wavelength (see Figure 2-9). By applying more energy (and considerable skill!) you may be able to create one wavelength or even 3/2 wavelengths. If you try this, note the correlation between higher energy and shorter wavelengths. A long piece of rubber tubing (about 12 feet long) from the laboratory is particularly well-suited for this demonstration.

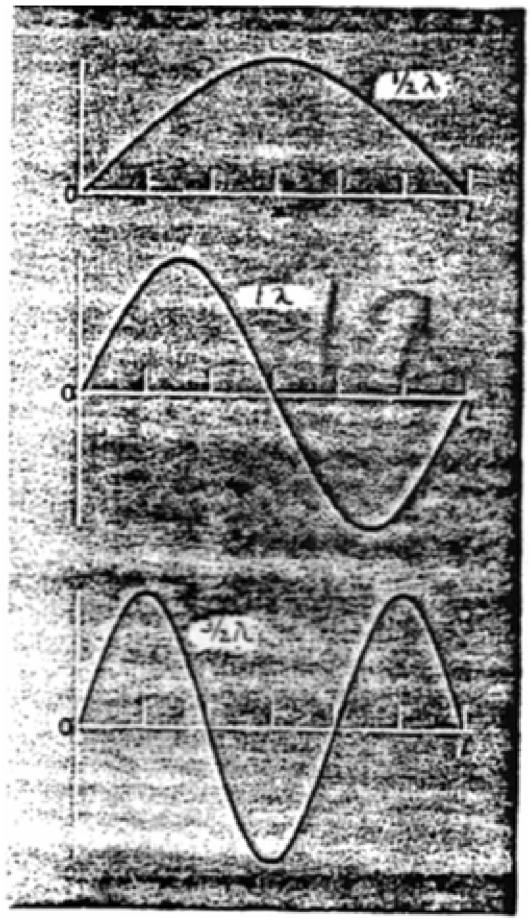


Fig. 2-9: Sketch of one-half wavelength (top), one wavelength (center), and 3/2 wavelength (bottom). Caption: *Standing waves from a jumping rope.*

The concept of electromagnetic radiation is an important tool in understanding many of the phenomena of everyday life. The microwaves in a microwave oven, for example, are electromagnetic radiation just like visible and infrared radiation, but of lower energy and longer wavelength. Radiowaves are longer yet. Each radio or television station has a characteristic frequency, which of course corresponds to a characteristic energy and wavelength. These radiowaves, ranging from about 0.1 meter to 1000 meters, are all around us, though they become evident only when an appliance that uses them to produce sound vibrations is turned on. Recently ELF radiation, or extra low frequency radiation, has become a center of controversy. Such very low energy radiation, created near electric power lines and even electric appliances, has always been considered harmless. Recent studies have suggested that extensive exposure to ELF radiation may be associated with some forms of cancer.

High-energy forms of radiation like x-rays and gamma rays have been recognized to be dangerous to the human body for decades, but the potential for harm was not known when they were first discovered. Chapter 3 tells the story of how mysterious "rays" were used to unlock some of the secrets of the atom.

CONCEPTS TO UNDERSTAND FROM CHAPTER 2

Matter has mass and occupies space.

All the matter in the universe is formed from the elements.

The atom is the smallest possible unit of an element.

Elements combine to produce compounds.

The smallest possible unit of a compound is called a molecule.

Atoms of each element have a characteristic mass.

Each element has a gram atomic mass that expresses the relative mass of its atoms in grams.

One gram atomic mass of any element contains 6.02×10^{23} atoms. This is called Avogadro's number.

Electromagnetic radiation is a form of energy.

Gamma rays, x-rays, ultraviolet radiation, visible radiation, microwave radiation, and radio waves are all forms of electromagnetic radiation.

The frequency of electromagnetic radiation is directly proportional to its energy.

The wavelength of electromagnetic radiation is inversely proportional to its energy.

FACTS TO LEARN FROM CHAPTER 2

The names and symbols of the first twenty elements

Among the philosophers who contributed to the ideas of the atom and the elements were:

Democritus, a Greek philosopher of the fifth century B.C., who proposed the concepts of atoms and the elements. The Greeks believed there were four elements: fire, air, earth, and water.

Lucretius, a Roman philosopher of the first century B.C., who supported the atomic theory and defended natural laws over superstition.

Scientists who contributed to the development of chemical theory included:

Robert Boyle, an English scientist who wrote *The Skeptical Chemist* in 1661, suggesting that the elements could be identified through experiments.

Antoine-Laurent Lavoisier,a French Scientist who wrote *An Elementary Treatise on Chemistry* in 1789, forming the foundation of modern chemistry.

John Dalton, an English chemist who formulated the concept of gram atomic mass in the early eighteenth century.

One gram atomic mass of any element contains 6.02×10^{23} atoms. This is called Avogadro's number.

SKILLS TO PRACTICE FROM CHAPTER 2

Using the problem-solving skills developed in Chapter 1, along with Avogadro's number, you should be able to:

calculate the number of atoms of an element if you are given the mass.

calculate the mass of an atom of any element in grams.

Using the decoding skills you have learned in Chapter 2, you should be able to:

Distinguish between elements and compounds

Identify the elements in a compound from the compound formula

Name ____________________ Date ___________

PROBLEMS TO SOLVE
USING CONCEPTS, FACTS, AND SKILLS FROM CHAPTER 2

2-1. Applying your chemical knowledge, answer the following questions about the four elements hypothesized by the Greeks: fire, air, earth, and water. Defend your answers by explaining your reasoning.

a. Are any of these considered pure elements today?

b. Are any of these considered pure compounds today?

c. Are all of these considered to be matter today?

2-2. Consulting the periodic table or a list of elements, identify any elements (<u>not</u> compounds) in the following list by circling them.

a. Aluminum

b. Brass

c. Copper

d. Iron oxide

e. Iron

f. Steel

g. Uranium

h. Salt

i. Sodium

j. Tellurium

2-3. Consulting the periodic table or a list of elements, identify any elements (not compounds) in the following list by circling them.

a. Ca b. DNA

c. He d. HI

e. H_2 f. C

g. Cu h. CPA

i. NaCl j. Cf

2-4. Identify the sources of sodium from this list on the label of a box of pancake mix by circling them:

Enriched flour and corn flour (with niacin, reduced iron, thiamine mononitrate, riboflavin), dextrose, leavening (sodium bicarbonate, sodium aluminum phosphate), buttermilk, whey, shortening, freshness preserved with BHA and citric acid, salt, egg yolk and egg whites.

2-5. According to a newspaper report, warplanes dropped red balloons to deflect heat-seeking surface-to-air missiles fired by guerilla defenders. Why were red balloons used? Would another color have done as well? Explain using your knowledge of the electromagnetic spectrum; look at Figure 6!

2-6. Was Lucretius a scientist? Explain. (You may want to review the discussion of the scientific method in Chapter 1.)

2-7. Do you think the alchemists were scientists? Explain.

2-8. Circle the one in each pair with higher energy:

a. Microwaves or infrared radiation?

b. Visible light or radiowaves? c. Ultraviolet light or x-rays?

2-9. Circle the one in each pair with longer wavelength:

a. Microwaves or infrared radiation?

b. X-rays or ultraviolet radiation?

c. UV-A or UV-B?

d. An AM radio station broadcasting at a frequency of 1030kHz or an FM radio station broadcasting at 102.5 mHz? (You will need to review your knowledge of the metric system for this one!)

2-10. Circle the one in each pair with higher frequency:

a. Microwaves or infrared radiation?

b. Visible light or ultraviolet radiation?

c. Visible light or infrared light?

2-11. Window glass transmits most light in the visible region but comparatively little in the ultraviolet region. Is it likely you will get a sunburn on a sunny day driving in a car with the windows rolled up? Explain.

2-12. Circle the one in each pair with higher energy:

a. Blue light with a wavelength of 480 nm or orange light with a wavelength of 610 nm?

b. Shortwave radio with a wavelength of 20 m or police band radio with a wavelength of 100 m?

c. Microwave with a wavelength of 1 centimeter or radar with a wavelength of 10 cm?

2-13. How many atoms in:

a. 1 gram atomic mass of carbon?

b. 12.0 g of carbon?

c. 24.0 g of carbon?

2-14. How many pineapples in Avogadro's number of pineapples?

2-15. How many atoms in:

a. 1 gram atomic mass of oxygen?

b. 16.0 g of oxygen?

c. 32.0 g of oxygen?

d. 8.0 g of oxygen?

2-16. How many grams in one atom of hydrogen?

2-17. If Lavoisier met Democritus and they compared their views of the nature of matter, what would they agree on? How would their viewpoints differ?

2-18. A university review panel in 1990 criticized the accounts of some scientists who reported that they had created the element helium through "cold fusion" experiments because, the report said, of 'language which refers to experiments which support a particular hypothesis as "successes" and those which tend to refute it as "failures."' Why is it inconsistent with the scientific method to do so?