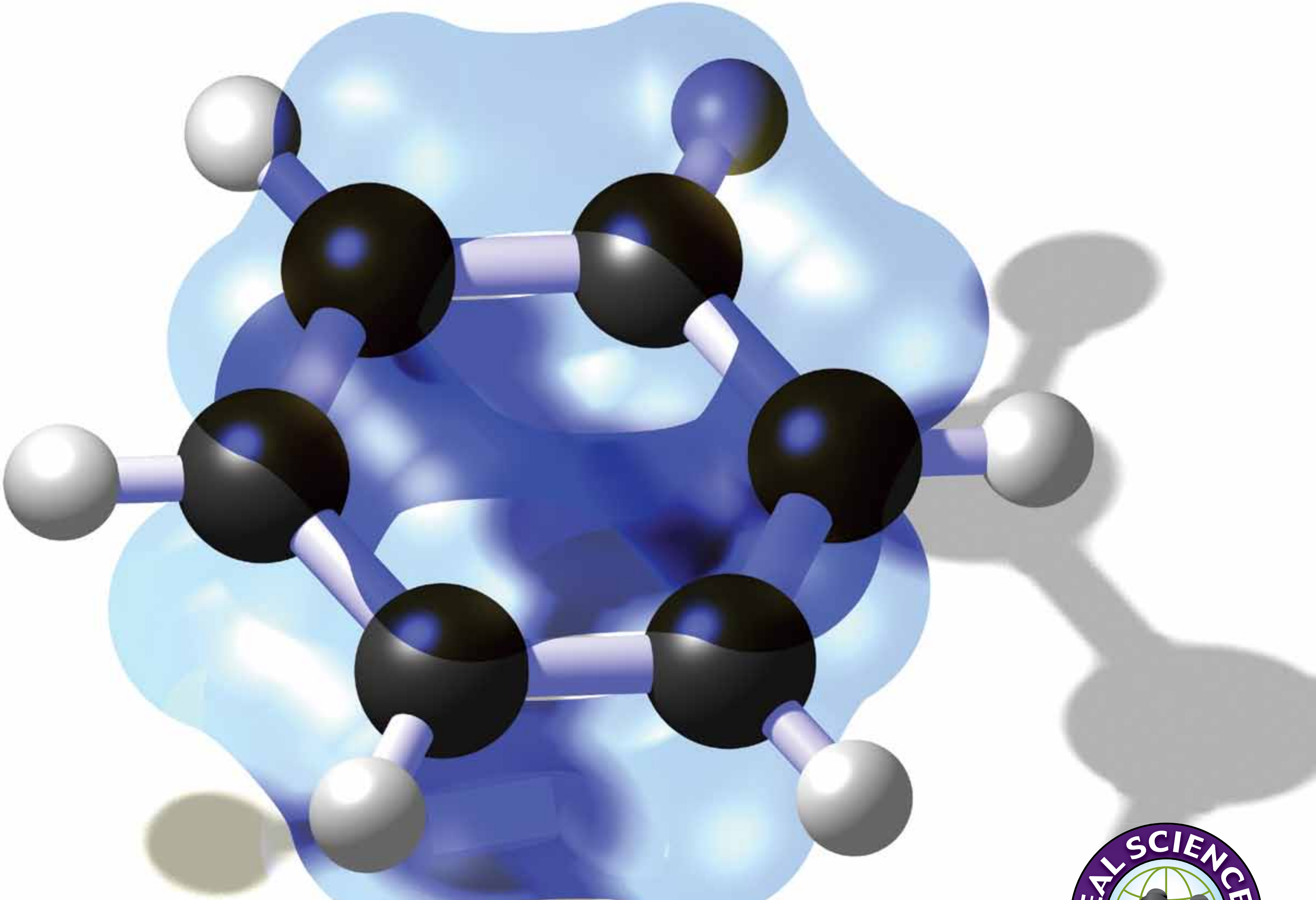
Grades 9-12

Focus On High School Chemistry



Rebecca W. Keller, PhD

Grades 9-12

Focus On High School Chemistry

Rebecca W. Keller, PhD

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Rebecca W. Keller, PhD

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The Periodic Table of Elements

Alkali metals | Alkaline earth metals | Metals | Non-metals | Metalloids | Synthetic elements (unconfirmed)

Transition metals

Halogens

Noble gases

1 H Hydrogen 1.0079																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.0122											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.0067	8 O Oxygen 15.9994	9 F Fluorine 18.9984	10 Ne Neon 20.1797
11 Na Sodium 22.9898	12 Mg Magnesium 24.3050											13 Al Aluminum 26.9815	14 Si Silicon 28.0855	15 P Phosphorus 30.9738	16 S Sulfur 32.066	17 Cl Chlorine 35.4527	18 Ar Argon 39.948
19 K Potassium 39.0983	20 Ca Calcium 40.078	21 Sc Scandium 44.9559	22 Ti Titanium 47.867	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.9380	26 Fe Iron 55.847	27 Co Cobalt 58.9332	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.64	33 As Arsenic 74.9216	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80
37 Rb Rubidium 85.4678	38 Sr Strontium 87.62	39 Y Yttrium 88.9059	40 Zr Zirconium 91.224	41 Nb Niobium 92.9064	42 Mo Molybdenum 95.96	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.9055	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.411	49 In Indium 114.82	50 Sn Tin 118.710	51 Sb Antimony 121.757	52 Te Tellurium 127.60	53 I Iodine 126.9045	54 Xe Xenon 131.29
55 Cs Cesium 132.9054	56 Ba Barium 137.327	57 *La Lanthanum 138.9055	72 Hf Hafnium 178.49	73 Ta Tantalum 180.9479	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.2	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.9666	80 Hg Mercury 200.59	81 Tl Thallium 204.3833	82 Pb Lead 207.2	83 Bi Bismuth 208.9804	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium 226.0254	89 **Ac Actinium 227.0278	104 Rf Rutherfordium 261	105 Db Dubnium 262.11	106 Sg Seaborgium 263	107 Bh Bohrium 264.12	108 Hs Hassium 269.13	109 Mt Meitnerium 266	110 Ds Darmstadtium (269)	111 Rg Roentgenium (272)	112 Cp Copernicum (277)	113 Uut Ununtrium (284)	114 Uuq Ununquadium (285)	115 Uup Ununpentium (288)	116 Uuh Ununhexium (289)	117 Uus Ununseptium	118 Uuo Ununoctium (293)

*The lanthinide series: Elements 58-71	58 Ce Cerium 140.116	59 Pr Praseodymium 140.9076	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.9253	66 Dy Dysprosium 162.50	67 Ho Holmium 164.9303	68 Er Erbium 167.26	69 Tm Thulium 168.9342	70 Yb Ytterbium 173.054	71 Lu Lutetium 174.967
**The actinide series: Elements 90-103	90 Th Thorium 232.0381	91 Pa Protactinium 231.0359	92 U Uranium 238.0289	93 Np Neptunium 237.0482	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (260)

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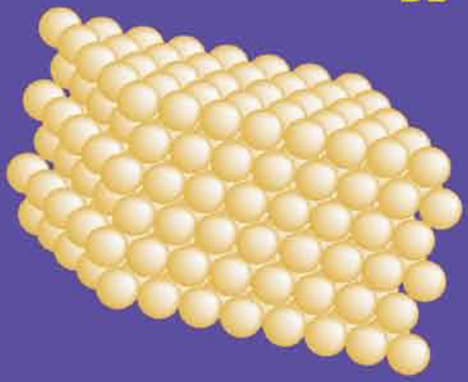
CONTENTS

CHAPTER 1 MATTER, MASS, AND MOLES



1.1 INTRODUCTION

1.1.1 Chemistry happens every day

Chemistry is the science of atoms and how they combine to form molecules. In many ways, it is the science that is most important to our everyday lives. Everything around us—the air, the ground, the chair we sit in, our clothes, and even our own bodies—is composed of atoms and molecules. Many of the things we do, such as running, swimming, jumping and eating involve chemical reactions. For example, when we eat a piece of chocolate or a bowl of pasta, chemical reactions inside the body convert these foods into energy. Most of that energy is used to keep our cells alive and to keep us warm, but we use some of the energy to play, think, breath, and even study chemistry! Chocolate contains fat and sugar molecules that can be chemically "burned" to produce useful energy. For example, molecules in your mouth called enzymes [en′-zīmz] begin breaking the food molecules into smaller pieces (proteins, fats, sugars). These smaller pieces are then used to make a host of other molecules (such as ATP—adenosine triphosphate [ə-de′-nə-sēn trī-fäs′-fāt]—and Acetyl CoA [ə-sēt′-əl kō–ā]) that enter complex metabolic pathways inside your cells. One such pathway is the citric acid cycle which uses Acetyl CoA to make the energy molecule ATP, the basic "fuel" used by all living things.

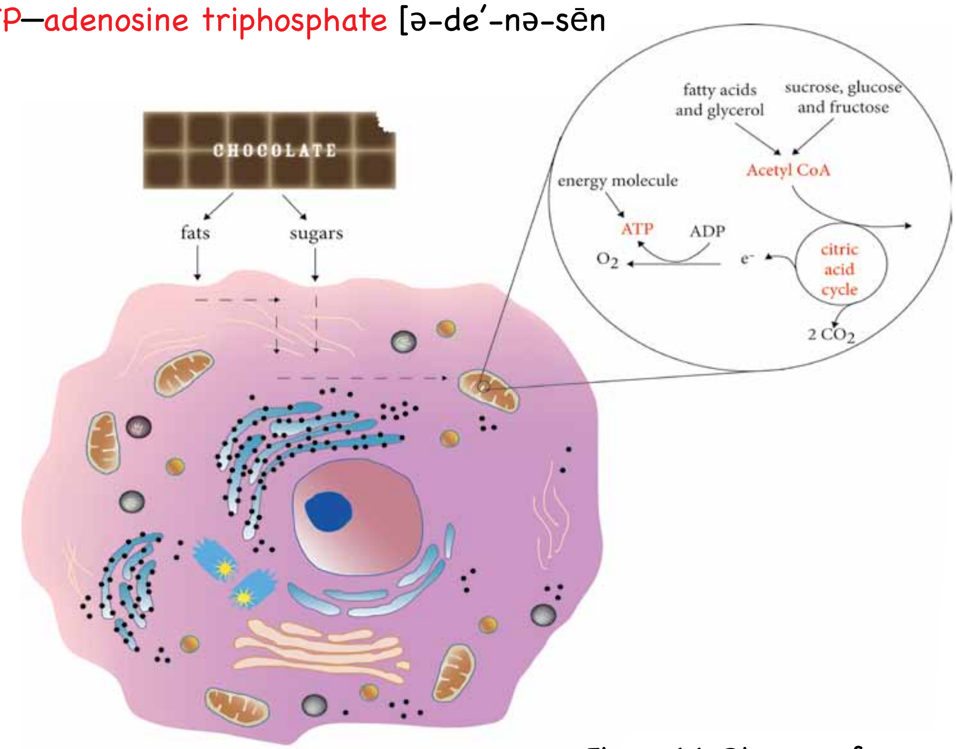


Figure 1.1: Diagram of a cell showing a simplified metabolic pathway.

1.1.2 What is chemistry?

Chemistry is the study of atoms and molecules, which are the smallest bits of matter that can undergo chemical reactions. At the simplest level, an atom is depicted as a little ball, and a molecule is shown as one or more atoms bonded together into a cluster of balls. A chemical reaction is a change in the way the atoms are bonded together. Chemistry is concerned with the properties of atoms and molecules and the way they react with each other.

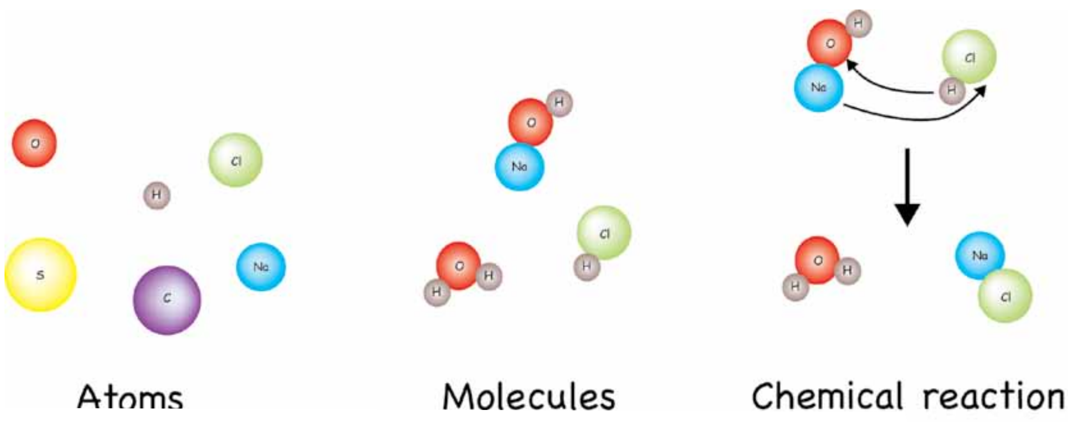


Figure 1.2: Atoms, molecules, and a chemical reaction.

1.1.3 Where did we get chemistry?

Today we know a great deal about the chemistry around us and inside our bodies, but it wasn't too long ago that chemistry was not very well understood. Where did the science of chemistry come from? How did we find out about chocolate, fats, sugars, enzymes, ATP, Acetyl CoA, acids, bases, water, proteins, carbohydrates, and everything else? The knowledge we have today about chemicals and their properties has accumulated over many centuries, beginning with the early Greek, Egyptian, and Chinese peoples. The word *chemistry* comes from the Greek word *chemeia*, which means "preparation of gold and silver." Much of modern chemistry has its roots in the work of artisans and alchemists of ancient Egypt, China, and Greece. Later in Europe, where the purification of gold and

silver was of prime importance, one aim of the early alchemists was to turn lead into gold. They experimented with different kinds of chemicals hoping that one day they would get rich! They never succeeded (do you know why?), but they did learn a great deal about chemicals and their properties.

Over time, the basic knowledge for the science we now call chemistry proved much more valuable than gold. Until the 1700s, chemistry was haphazard at best, but beginning with the investigations of Antoine Lavoisier, Joseph Priestly, and Robert Boyle, chemistry slowly developed into a powerful systematic science. The most important early experiments taught chemists how to separate mixtures from pure substances and how to weigh and measure the products of chemical reactions.

Today chemistry is divided into at least four main areas: analytical chemistry, synthetic chemistry, physical chemistry, and biochemistry. Analytical chemistry deals with what kinds of chemicals compose various substances. The word *analytical* comes from the Greek word *analeuin* which means "to resolve." Analytical chemists can determine, or resolve, the chemical makeup of a number of different and complex samples such as contaminated dirt, tissue samples from patients, or cloth and blood samples from a crime scene.

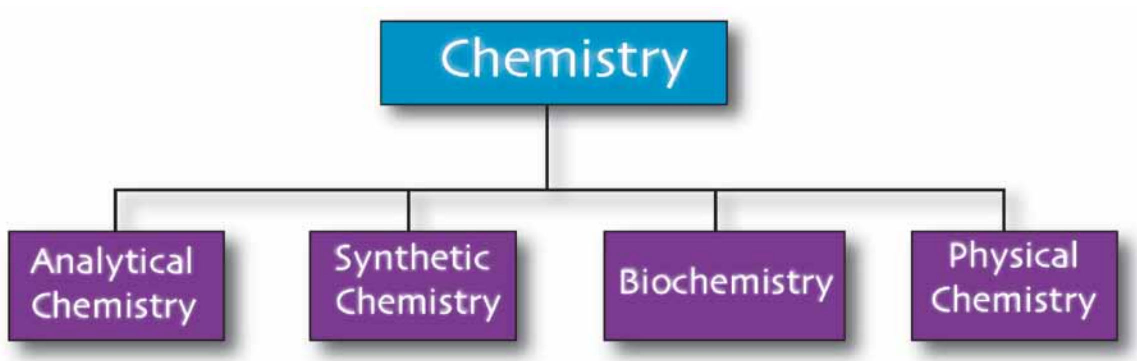


Figure 1.3: The major divisions of chemistry.

Synthetic chemistry deals with making new kinds of chemicals or chemicals not found naturally. The word *synthetic* comes from the Greek prefix *syn* which means "with or together," and the Greek word *thesis*, which means "to place or put." So synthetic chemists use their knowledge of how molecules react with each other to *synthesize* or put together new molecules. Nylon is a fiber used in clothing that was designed by synthetic chemists.

Physical chemistry focuses on how chemical reactions and other properties arise from the basic laws of physics. The word *physics* comes from the Greek word *physis*, which means "nature or natural growth." So physical chemistry is concerned with how natural forces are involved in molecular bonding, for example, or how some chemical reactions give off or absorb heat. Biochemistry is concerned with the complicated chemistry that takes place inside living things. The prefix *bio-* comes from the Greek word *bios*, which means "life," so biochemists study the chemistry of life. For example, biochemists have figured out how chocolate is turned into energy!

1.2 ATOMS

1.2.1 Protons, neutrons, and electrons

Atoms are the smallest distinctive chemical units of matter. They are made of three smaller particles: protons [prō′-tonz], neutrons, [nü′-tronz] and electrons [i-lek′-tränz]. Protons and neutrons make up the center of an atom, called the atomic nucleus. Even though the nucleus is very tiny (it would be invisibly small if it were shown at its real size in Figure 1.4), it contains essentially all the *mass* of the atom. Mass is different from weight and is a property that makes matter resist being moved (see section 1.2.2). The electrons occupy the space surrounding the nucleus, called the electron cloud. Even though the cloud is large (compared to the nucleus), it has almost no mass. Figure 1.4 shows a simple electron cloud, but the electron cloud is really made up of several smaller clouds, called orbitals. Some orbitals have odd shapes; they are not simple and round as shown in the diagram. We will learn more about the different kinds of orbitals in Chapter 2 (see Section 2.2).

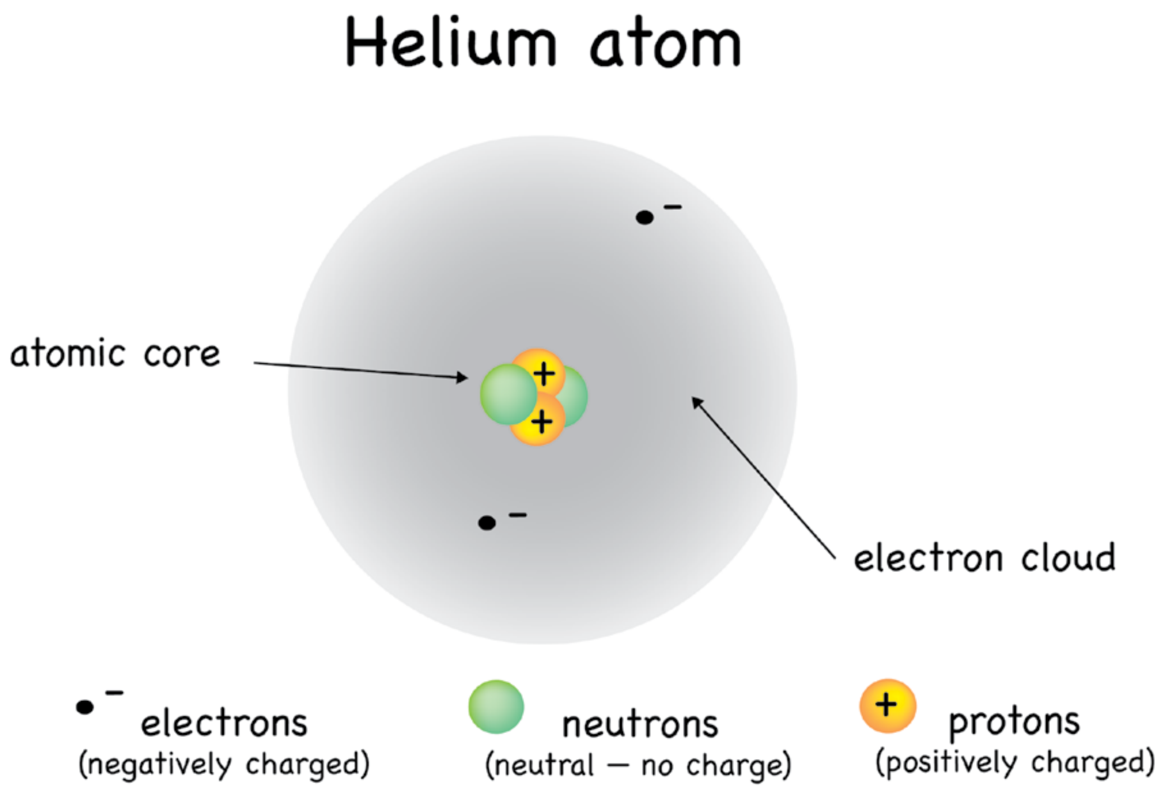


Figure 1.4: A helium atom showing the atomic core (with two protons and two neutrons) and the electron cloud (with two electrons).

Every proton in the nucleus of an atom carries one unit of *positive* electric charge, and every electron carries one unit of *negative* electric charge. Neutrons have no electric charge. (They're neutral, that's why they're called neutrons!) The negative charge of one electron can cancel the positive charge of one proton, and *vice versa*. Every atom must have a zero electric charge overall. Therefore...

the number of electrons equals the number of protons.

For example, a hydrogen atom has 1 electron and 1 proton, a carbon atom, which has 6 protons, also has 6 electrons and a gold atom has 79 protons and 79 electrons.

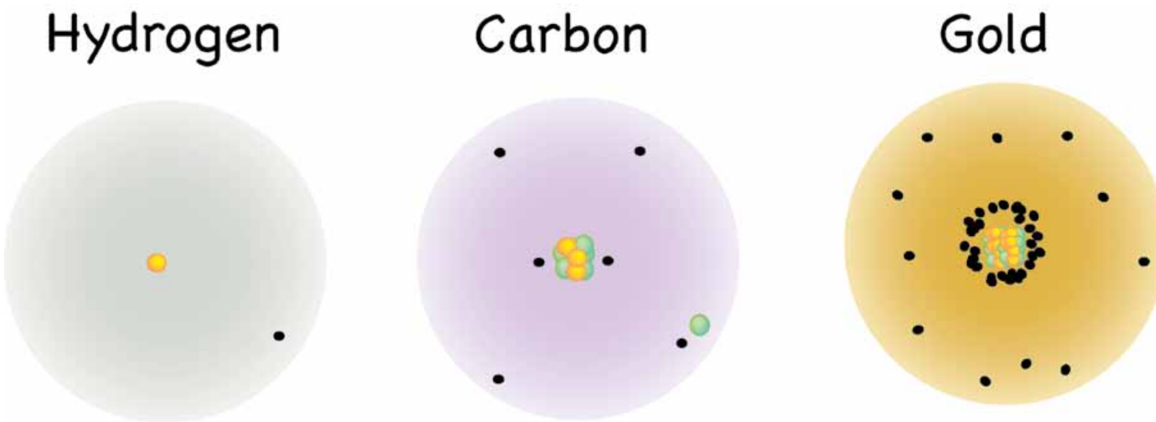


Figure 1.5: A hydrogen atom with 1 proton and 1 electron, a carbon atom with 6 protons and 6 electrons, and a gold atom with 79 protons and 79 electrons.

1.2.2 Matter and mass

Atoms are one of the most basic forms of matter. Tangerines and automobiles, baseballs and bowling balls, and even the wind and the rain are all made of matter. Everything we can see, touch, taste, or smell is made of matter. But what *is* matter?

Matter is a general term that describes anything that occupies space and has mass. Recall that mass is different than *weight*. Mass is a property that makes matter resist being moved (*intertia*), but weight is a force caused by the earth's gravity. For example, a bowling ball and a feather both have mass but the bowling ball has more mass than the feather. On earth, the bowling ball will *weigh* more than the feather, but in space neither the bowling ball nor the feather weighs anything at all. Now think about taking a bowling ball and shaking it back and forth. It is harder to shake than a feather. Even in space, where neither the bowling ball nor the feather has weight, the bowling ball is still harder to shake. This is because the bowling ball has more mass than the feather. It resists being moved, and it has more *inertia* than the feather.

1.2.3 Atomic mass and atomic weight

Since atoms are a form of matter, they must have mass. We can estimate the mass of an atom just by counting its protons and neutrons. The masses of atoms are most conveniently measured in atomic mass units, or amu.

Each proton and each neutron has a mass of 1 amu.

An electron also has mass, but it's only about 1/2000 the mass of a proton or neutron, so its mass is usually ignored in computing atomic mass. A hydrogen atom, with 1 proton and no neutrons, has a mass of 1 amu; a carbon atom, with 6 protons and 6 neutrons, has a mass of about 12 amu; and a gold atom, with 79 protons and 118 neutrons, has a mass of 197 amu. These values are only estimates. The true atomic mass is always a bit different than just the sum of protons and neutrons, but it is usually close.

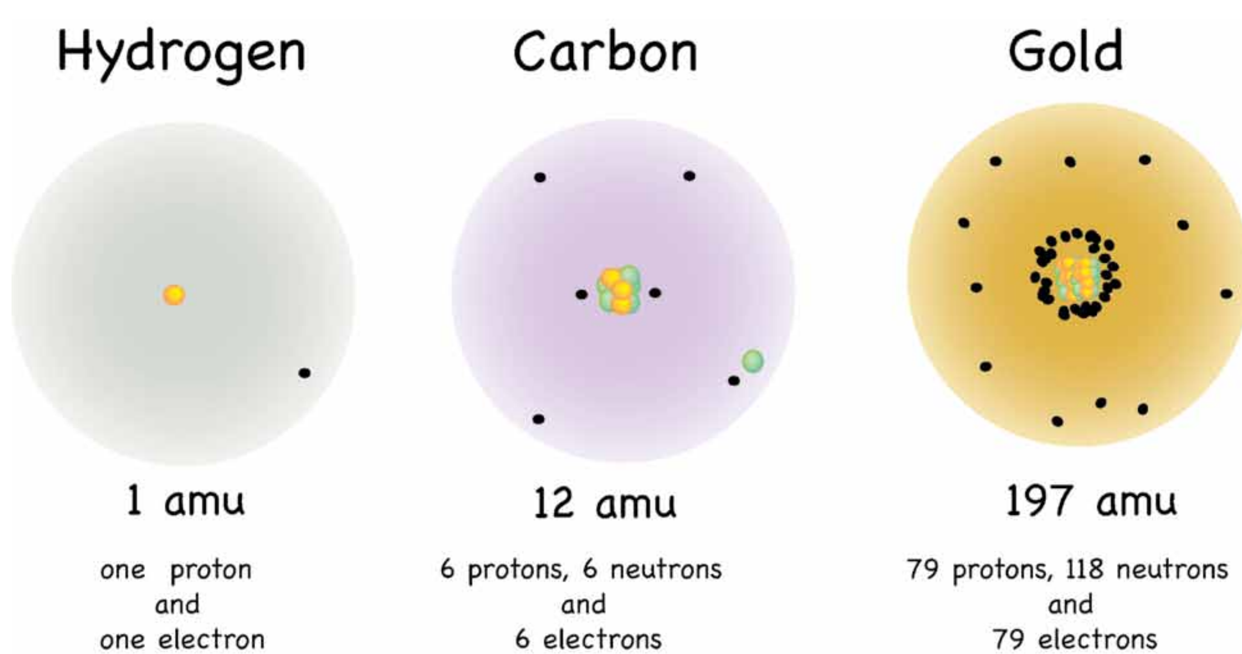


Figure 1.6: A hydrogen atom with an atomic mass of 1 amu, a carbon atom with an atomic mass of 12 amu, and a gold atom with an atomic mass of 197 amu.

Mass and weight are different, but on Earth mass always gives rise to weight (the force due to gravity acting on mass). So we often talk about the weight of something when we really mean mass. Very often chemists speak of "atomic weight," but this just means the same thing as atomic mass. The atomic mass (or weight) of each element is listed on the periodic table. For example, the atomic mass of oxygen (O) is listed on your table as 15.9994 amu, and the atomic mass of iron (Fe) is listed as 55.847 amu.

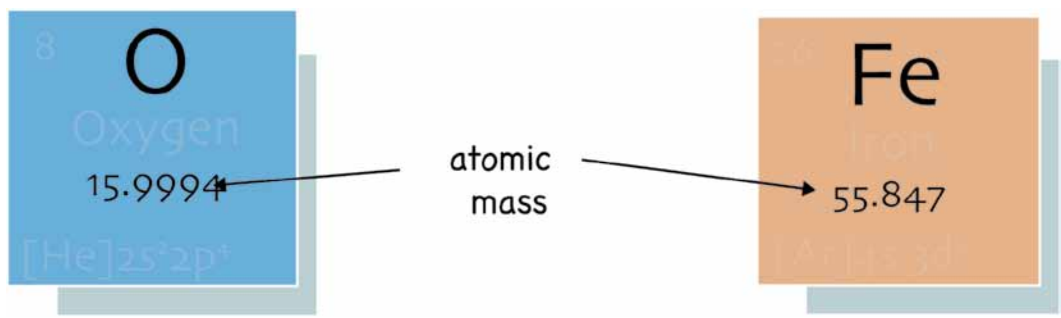


Figure 1.7: The element oxygen with an atomic mass of 15.9994 amu and the element iron with an atomic mass of 55.847 amu as represented on the periodic table.

Atoms are very small, so the mass of one atom is also very tiny. However, if there are enough atoms together they can weigh a lot. *The weight (or mass) of every object around us is just the sum of the masses of all the atoms in it.* A bowling ball has a lot of mass, but it all comes from the masses of the atoms inside. Moreover, if the mass is doubled, the number of atoms is also doubled, as is the weight.

1.2.4 Moles: counting and weighing atoms

The fact that the weight of an object is the sum of atomic masses has one very important consequence: You can count atoms by weighing them! Atoms are *much* too small to count one by one, so chemists count them in big groups, called moles. A mole is defined such that 1 mole of carbon atoms, each with a mass of 12 amu, weighs 12 grams. In general:

A mole of atoms weighs the same (in grams) as its atomic mass (in amu).

For example, a helium atom has a mass of 4 amu, so a mole of helium atoms weighs 4 grams. Two moles of helium atoms weighs 8 grams, three moles weighs 12 grams, and so on. (*How much would half a mole of helium atoms weigh? How many moles are there in 16 grams of helium?*)

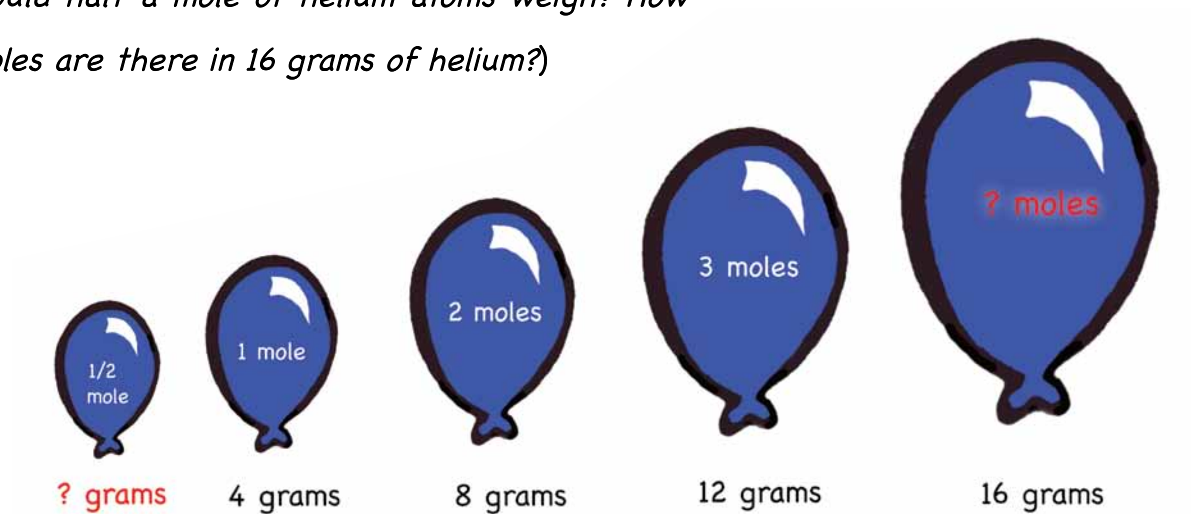


Figure 1.8: Grams of helium increase as the number of moles increases. Since helium has a mass of 4 amu, the number of grams increases by a multiple of 4 for every mole of helium.

An oxygen atom has a mass of 16 amu; a mole of oxygen atoms weighs 16 grams. (How much would two moles of oxygen atoms weigh? How many moles are there in 4 grams of oxygen?)

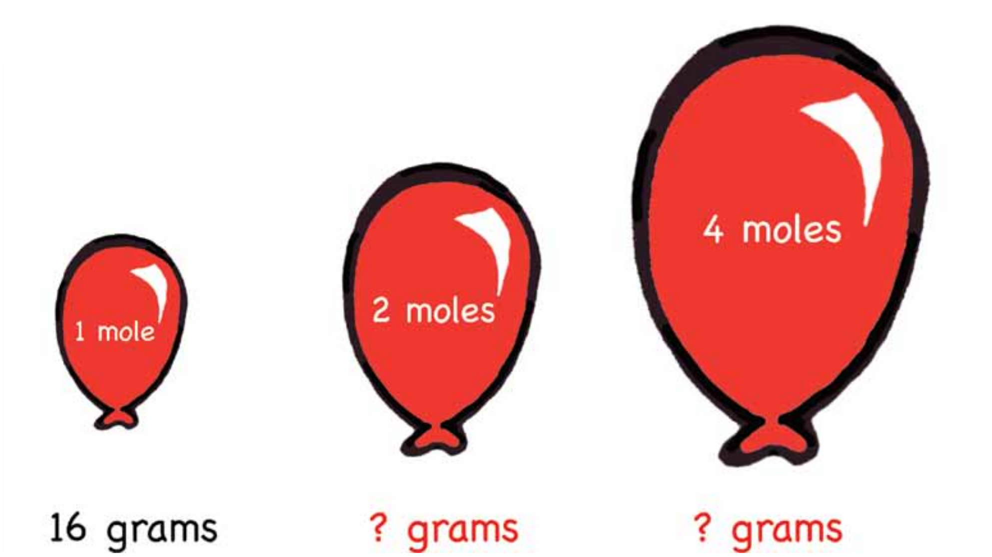


Figure 1.9: Grams of oxygen increase as the number of moles increases. Since oxygen has a mass of 16 amu, the number of grams increases by a multiple of 16 for every mole of oxygen.

A gold atom has a mass of 197 amu; a mole of gold atoms weighs 197 grams. (How much would two moles of gold weigh? four moles?)

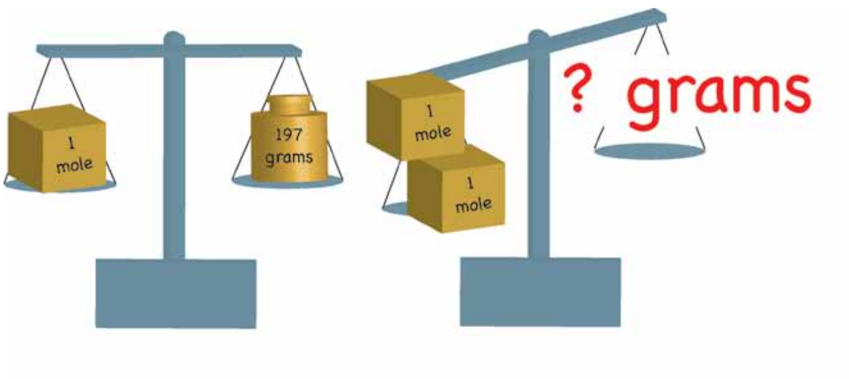
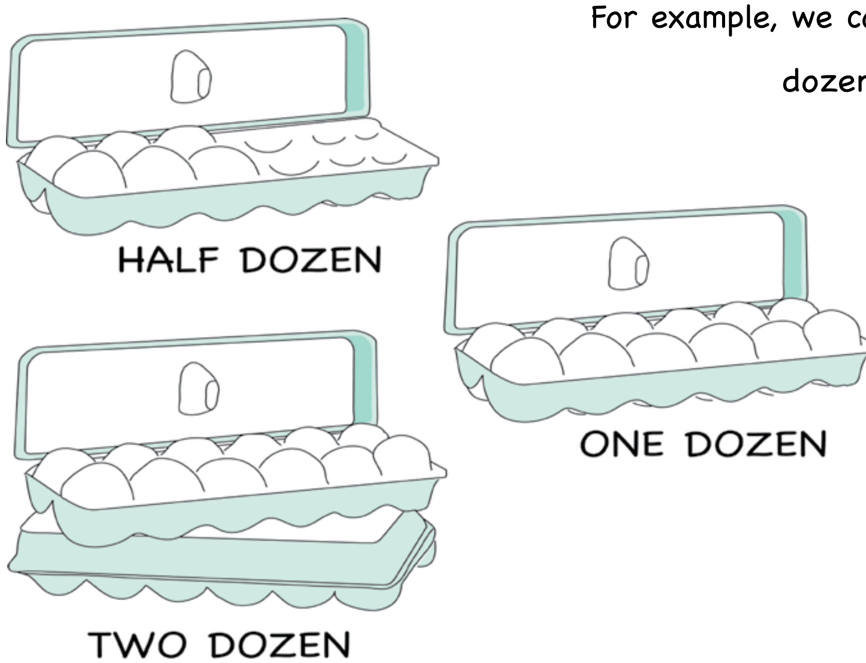


Figure 1.10: The number of moles is balanced by quantity in grams.

A mole is just a name that represents a certain number of atoms, molecules, ions, or even umbrellas. We count things all the time by using names that represent certain amounts. For example, we count 12 eggs by calling them a dozen. Twenty-four eggs are two dozen; six eggs are a half dozen, and so on. A dozen is just a group of 12, so when you count by dozens, you are counting by groups of 12. In the same way, chemists count by groups called moles.



The only difference is that a mole is a VERY large number:

1 mole = 602,200,000,000,000,000,000,000

(If you are familiar with scientific notation, a mole can also be written in a more convenient form: 1 mole = 6.022×10^{23}.) A mole is 6022 followed by 20 zeros. That's such a big number, it won't even fit on most calculators! It's so big that if you had a mole of marbles, it would be bigger than the moon! But atoms are very tiny, so a mole of atoms is a nice, manageable size. A mole of most atoms will fit in the palm of your hand.

It is important to note that one mole of any atom is the same number of atoms, no matter how much the sample weighs. One mole of carbon atoms has the same number of atoms as 1 mole of gold atoms, even though the carbon weighs 12 grams and the gold weighs 197 grams.

The mole is a very important concept in chemistry. It allows chemists to tell how many atoms or molecules they have in their samples just by weighing them. For a pure element, if you know the weight of your sample in grams, then you can find the number of moles in your sample by dividing by the atomic mass.

$$\text{moles} = \frac{\text{weight in grams}}{\text{atomic mass in amu}}$$

Without the concept of the mole, it would be difficult to tell how many atoms you have in a sample, and much of chemistry would not be possible.

1.3 THE PERIODIC TABLE

1.3.1 Introduction

All of the elements are organized into a table called the periodic table of elements. The periodic table lists all the elements and gives some of their properties, such as number of protons, atomic weight, and chemical symbol. More importantly, the periodic table organizes the elements into families according to such chemical properties. By noting where an element is in the periodic table, you can often tell how it will behave when mixed with other elements.

1.3.2 The elements

As we saw in Section 1.2, the nucleus of every atom is made of protons and neutrons, and some nuclei have more protons than others. For example, an atom with only one proton and one electron is a hydrogen atom, but an atom with two protons, two neutrons, and two electrons is a helium atom. An atom with three protons, three neutrons, and three electrons is a lithium atom.

We call the different kinds of atoms elements. There are 92 elements found naturally

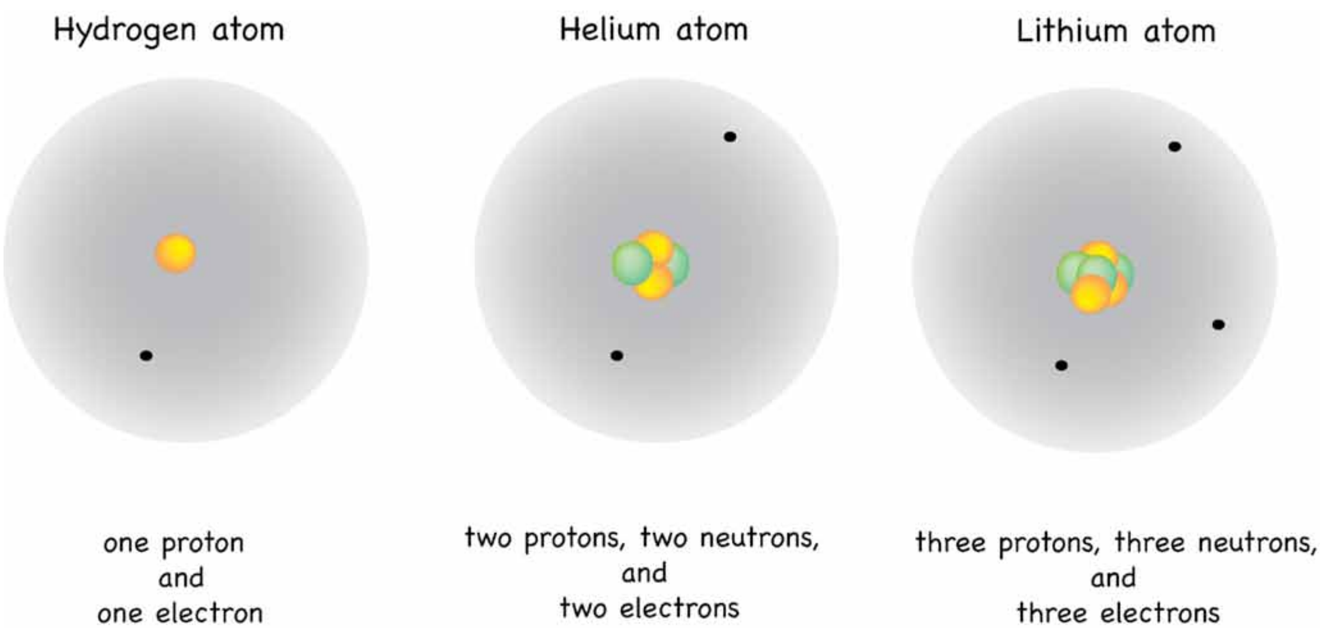


Figure 1.11: A hydrogen atom with one proton and one electron, a helium atom showing two protons, two neutrons, and two electrons, and a lithium atom with three protons, three neutrons, and three electrons.

and about a dozen additional elements that have been synthesized artificially by humans. The periodic table lists all the different elements, that is, all the different possible kinds of atoms. Hydrogen, helium, carbon, oxygen, gold, and plutonium are all elements.

1.3.3 Mixtures and pure substances

Matter—which is everything you can see, taste, and touch—is made of different kinds of atoms. Matter can be classified into two categories: mixtures and pure substances. We will learn more about mixtures in Chapters 6 and 7.

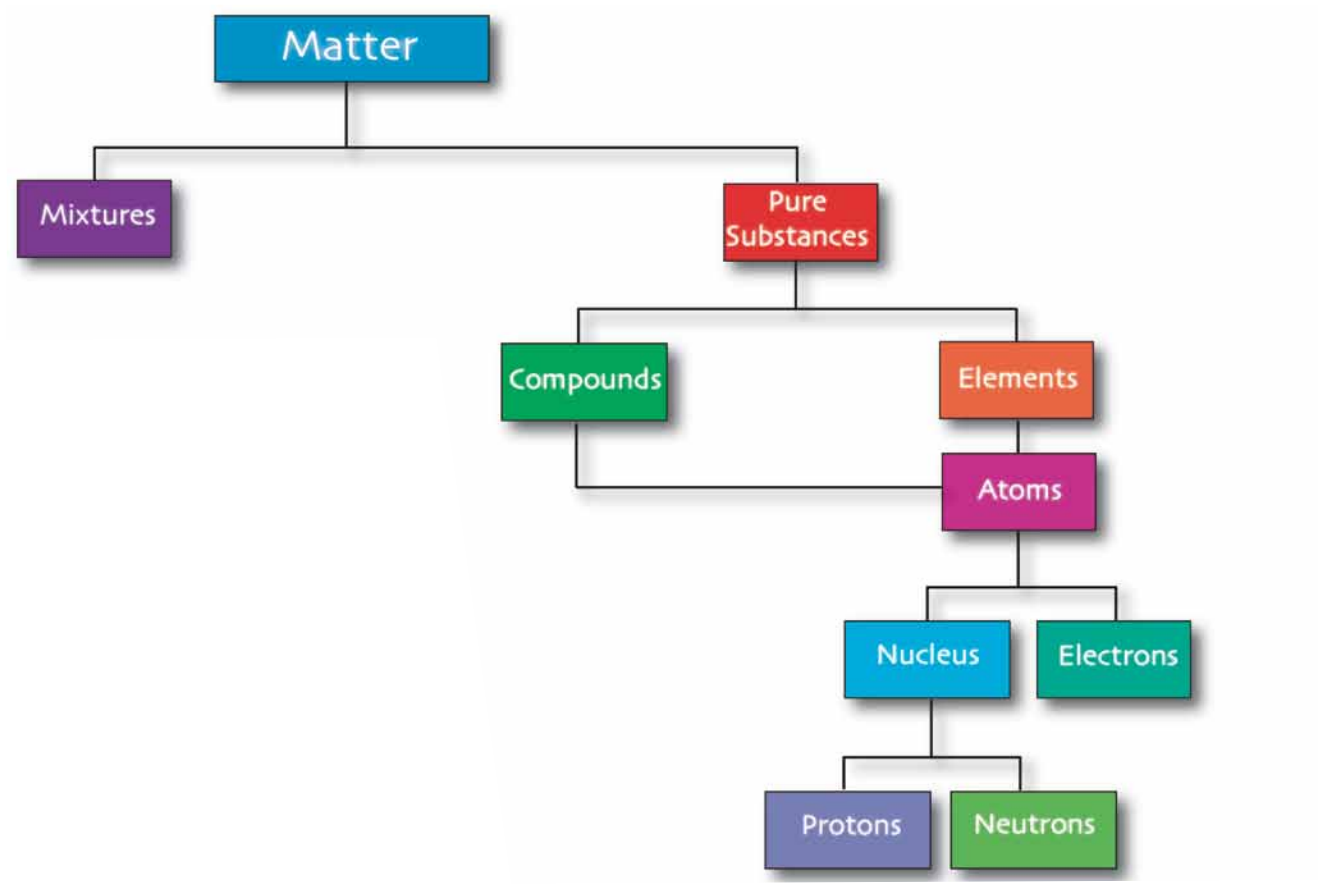


Figure 1.12: Matter can be divided into two categories; mixtures and pure substances. Pure substances can be further divided into compounds and elements. Elements are composed of atoms, which contain a nucleus (with the protons and neutrons) and the electrons.

A pure substance can be either a *compound* or an *element*. Some pure substances are made of only one kind of atom. For example, pure gold contains *only* gold atoms and nothing else. Likewise, pure graphite, such as the graphite in your pencil, contains *only* carbon atoms and nothing else. All pure metals such as aluminum, iron, and copper contain *only* one kind of atom. Pure oxygen gas is made of *only* oxygen atoms, and pure nitrogen gas is made of *only* nitrogen atoms.

We call such substances elemental, as in *elemental gold*, or *elemental carbon*, to indicate that they are composed of only one kind of atom.

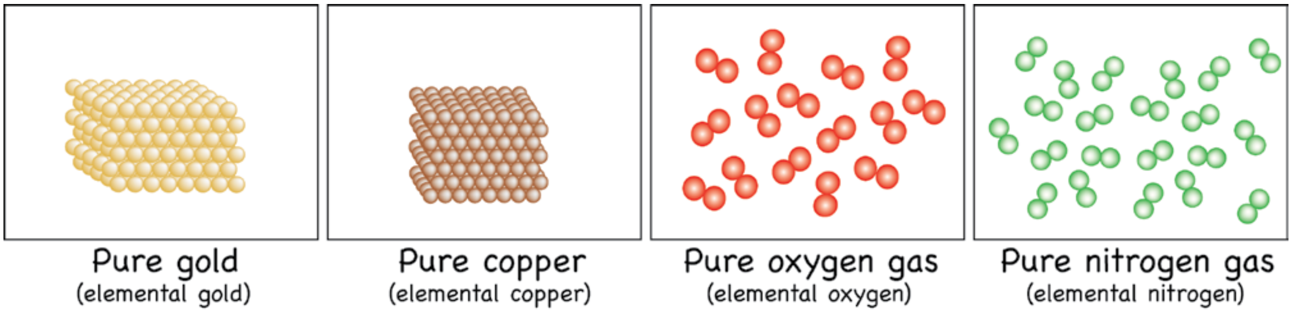


Figure 1.13: Pure gold, pure copper, pure oxygen gas and pure nitrogen gas are all elemental substances.

Other pure substances contain more than one element but are composed of only one type of molecule. Pure water, for example, is made of two different elements: hydrogen and oxygen. Pure ammonia contains three hydrogen atoms and one nitrogen atom, and pure carbon dioxide contains two oxygen atoms and one carbon atom. Even though they contain more than one kind of atom, they are considered to be pure substances because they are composed of only one kind of molecule. We call such pure substances compounds.

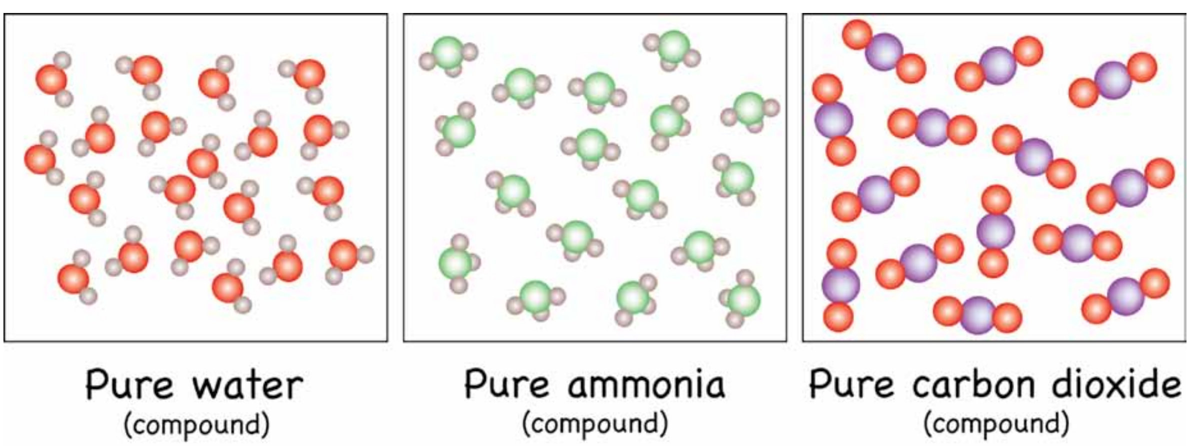


Figure 1.14: Pure water, pure ammonia and pure carbon dioxide are all compounds.

In general, a compound is *two or more atoms bonded together in a fixed ratio.* (For example, there is always one oxygen atom to two hydrogen atoms in a water molecule for a ratio of one to two.) Because the atoms in compounds have a fixed ratio, they are also considered pure substances.

Matter can also exist in mixtures. A mixture is defined as two or more substances physically (but not chemically) combined. The air we breathe is a mixture of nitrogen gas, oxygen gas, and other trace substances. Tap water is a mixture of water molecules and small amounts of metals and chlorine. Unlike a compound, the ratio of the components of a mixture are not fixed. Therefore, mixtures are *not* pure substances.

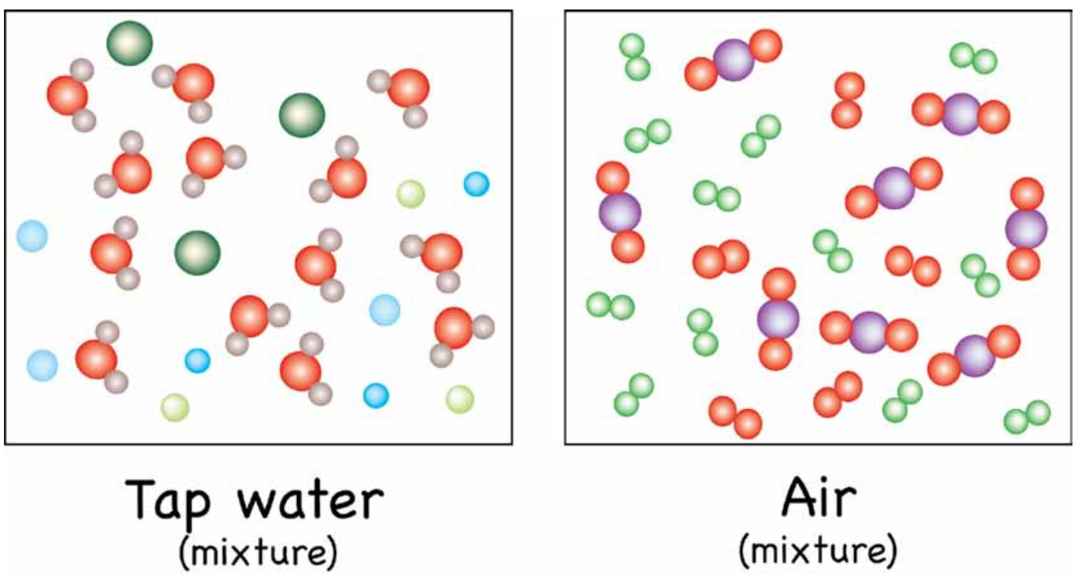


Figure 1.15: Tap water is a mixture of water and other ions. Air is a mixture of different gases.

1.3.4 Where did we get the periodic table?

At one point in history, many elements were known, but there was no suitable way to organize them. This problem was largely solved by a Russian chemist named Dmitri Mendeleev [dmē'-trē men-də-lā'-əf]. Mendeleev loved playing card games, and one day he decided to put all of the elements known to him on individual playing cards.

Then, he organized the cards according to their atomic mass and chemical properties. When he did this, he discovered that elements with similar chemical properties fell into a particular pattern or periodicity. He wasn't sure what to do with hydrogen, the lightest element, so originally he left it out. He started with lithium (Li), and noted that elements, such as sodium (Na) and potassium (K), with chemical properties similar to lithium, were spaced eight elements apart: Li is element number 3 (3 protons, 3 electrons); Na is element number 11 (11 protons, 11 electrons); K is element number 19 (19 protons, 19 electrons), and so on. He lined them up with increasing atomic weight and then grouped them into families with similar chemical properties. An early periodic chart by Dmitri Mendeleev looked something like this:

PERIODIC TABLE: DMITRI MENDELEEV (1834 — 1907)

SERIES	GROUPS O	I	II	III	IV	V	VI	VII	VIII
1		H = 1.008							
2	He = 4.0	Li = 7.03	Be = 9.1	B = 11.0	C = 12.0	N = 14.04	0 = 16.00	F = 19.0	
3	Ne = 19.9	Na = 23.05	Mg = 24.3	Al = 27.0	Si = 28.4	P = 31.0	S = 32.06	Cl = 35.45	
4	Ar = 38	K = 39.1	Ca = 40.1	Sc= 44.1	Ti = 48.1	V = 51.4	Cr = 52.1	Mn = 55	Fe = 55.9, Co = 59
5		Cu = 63.9	Zn = 65.4	Ga = 70	Ge = 72.3	As = 75	Se = 79	Br = 79.95	Ni = 59, (Cu)
6	Kr = 81.8	Rb = 85.4	Sr = 87.6	Y = 89	Zr = 90.6	Nb = 94.0	Mo = 96	(-)	
7		Ag = 107.9	Cd = 112.4	In = 114.0	Sn = 119	Sb = 120	Te = 127	I = 127	Ru = 101.7, Rh = 103
8	Xe = 128	Cs= 132.9	Ba = 137.4	La = 139	Ce = 140	(-)	(-)	(-)	Pd = 106.5, (Ag)
9		(-)	(-)	(-)	(-)	(-)	(-)	(-)	
10	(-)	(-)	(-)	Yb= 173	(-)	Ta = 183	W = 184	(-)	Os = 191, Ir = 193
11		Au= 197.2	Hg = 200.0	Tl = 204.1	Pb = 206.9	Bi = 208	(-)	(-)	Pt = 194.9, (Au)
12	(-)	(-)	Rd= 224	(-)	Th = 232	(-)	U = 239		
	R	R_2O	RO	R_2O_2	RO_2	R_2O_6	RO_3	R_3O_7	RO_4

(1871) ADAPTED TABLE FROM ANNALEN, SUPPL. VIII, 133 (1871); REVISED 1898 AFTER THE DISCOVERY OF RADIUM (Rd, MODERN Ra)

Figure 1.16: An early periodic table. [Redrawn by R.W. Keller from several composite tables.]

1.3.5 The periodic table today

Mendeleev's chart had gaps in the pattern that suggested there should be other elements. So even though he didn't know all of the elements we know today, he was able to predict new elements that would fit into the empty spaces on his chart. He could also tell roughly what atomic mass they would have.

Today, we know all of the naturally occurring elements, and scientists have even been successful in synthesizing many elements that do not occur naturally. The modern periodic table is organized in three ways. First, *each element is identified by the number of protons alone*; the number of neutrons doesn't matter. The number of protons is called the atomic number. The atomic number determines the element. Atomic number 1 (1 proton) is hydrogen, atomic number 2 (2 protons) is helium, atomic number 6 (6 protons) is carbon, atomic number 79 (79 protons) is gold, and so on.

Second, *elements increase in atomic number from left to right across the rows*. The first row has just hydrogen (atomic number 1) on the left side and helium (atomic number 2) on the right side. The second row elements include lithium (atomic number 3) in the left-hand column through neon (atomic number 10) in the extreme right-hand column.

The third row elements are sodium (element 11) through argon (element 18), and so on. Each row is called a period (hence the name "periodic table").

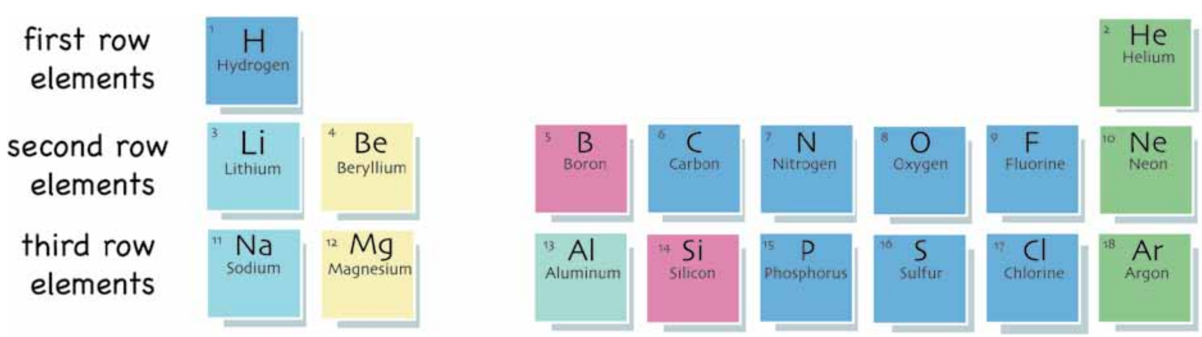


Figure 1.17: The first three rows of the periodic table.

Third, *elements with similar chemical properties are lined up in the same column.* By chemical properties chemists mean "the way the atoms react with each other to form compounds." For example, the elements in the left-hand column (lithium, sodium, potassium, rubidium, cesium, and francium) are called the alkali metals. They are all soft, white metals, and they all react strongly with water. Sodium hisses and fumes when tossed into water, and potassium explodes. It can even cause fires from moisture in the air.

Rubidium and cesium are similar. Similarities in physical properties (like color, hardness, and metallic character) and chemical properties (like reacting with water in similar ways) are examples of the kinds of properties Mendeleev used to create his early periodic tables.

The elements in the next-to-last column (fluorine, chlorine, bromine, iodine, and astatine) are called the halogens. Most are strongly colored and melt or vaporize at low temperatures. Fluorine is a colorless gas, chlorine is a yellow-green gas, bromine is a dark red-brown liquid, and iodine is a dark brown solid. The halogens react with most metals, but they react especially violently with the alkali metals. The reaction between any of the alkali metals and any of the halogens always creates a saltlike compound. For example, sodium and chlorine react to

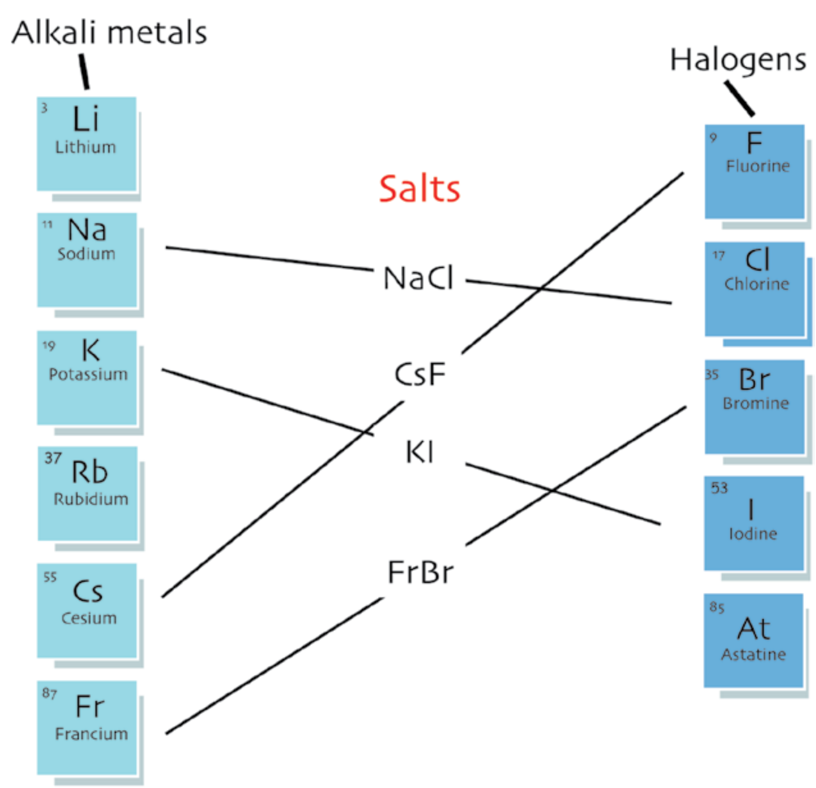


Figure 1.18: The alkali metals and the halogens react to form salts.

form sodium chloride, which is ordinary table salt. Potassium and chlorine form potassium chloride, which is sometimes called "light salt" and is so similar to table salt that you can eat it, and it almost tastes the same. Similarly, lithium reacts with bromine, cesium reacts with fluorine, and so on. All these combinations create saltlike compounds.

The very last column (helium, neon, argon, krypton, xenon, radon) contains the inert gases, or noble gases. As their name implies, they are all gases, and they are all very unreactive (with any element). Only a very few compounds containing the noble gases have ever been discovered.

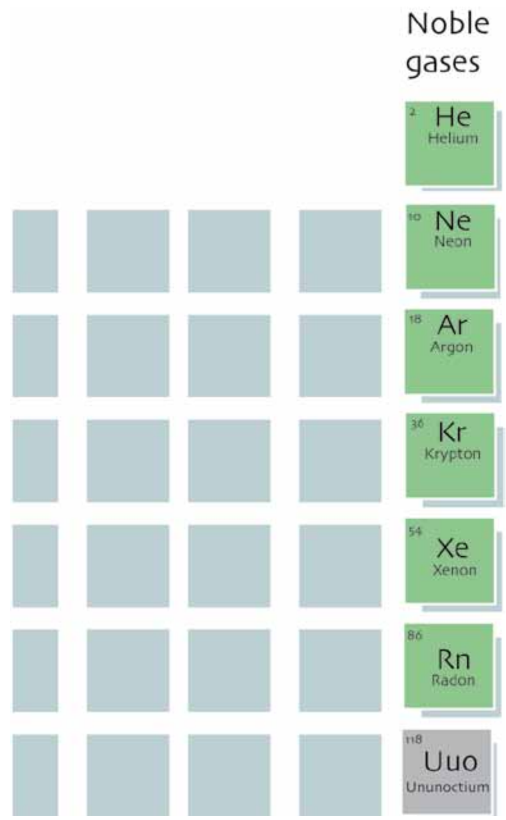


Figure 1.19: The noble gases.

Finally, the periodic table is also divided into two large blocks: the main group and the transition metals (also called transition elements). The main group elements are located on the right-hand and left-hand sides of the periodic table.

The difference between the main group and the transition metals is in the way their electrons are organized. To jump ahead a bit, the main group elements have their outermost electrons (the ones that are responsible for most chemical properties) in *s* and *p* shells, while the transition metals have their outermost electrons in *d* shells. We will learn more about electron shells and how the electrons in atoms are organized in Chapter 2.

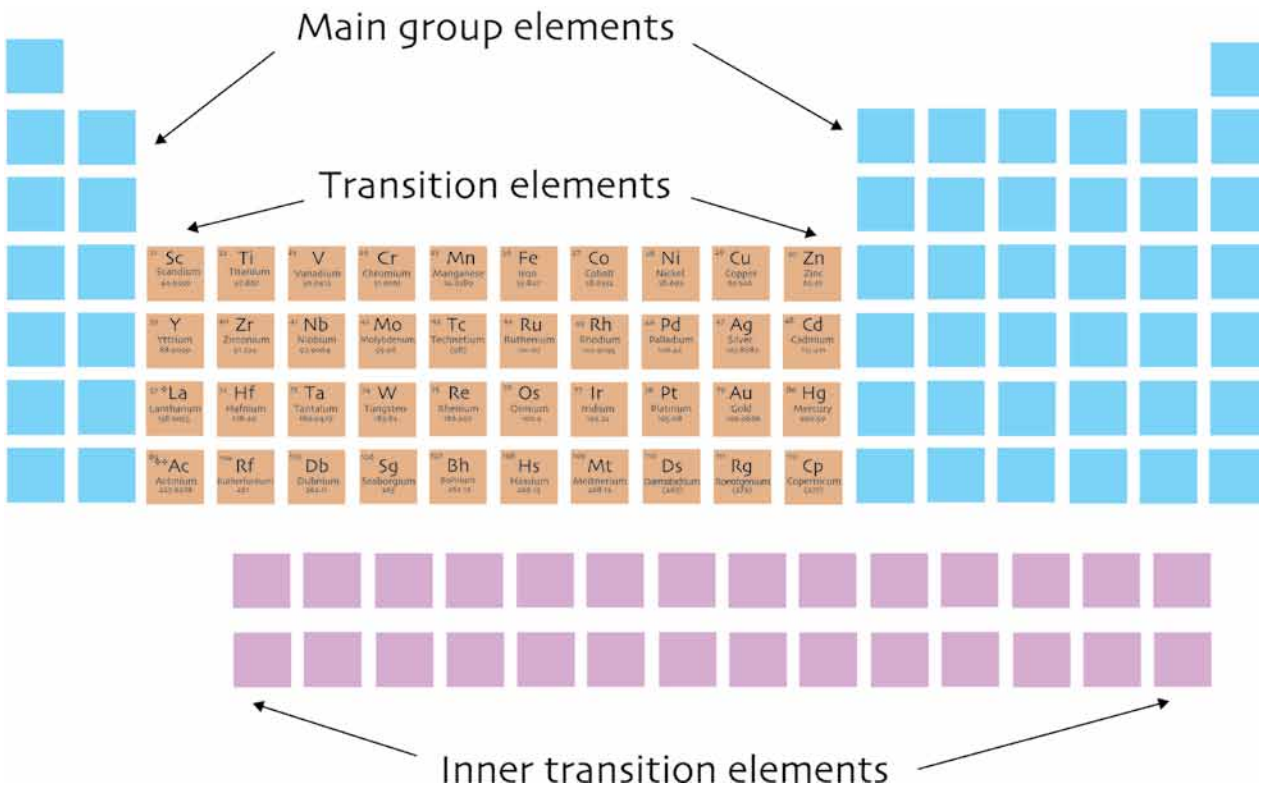


Figure 1.20: The main group elements (blue), the transition elements (orange), and the inner transition elements (pink).

1.4 SUMMARY

- Everything around us is made of atoms. Chemistry is the study of how atoms combine to make molecules and how both atoms and molecules react with each other to form new substances.
- Every atom is made of three basic particles: protons, neutrons, and electrons. The protons and neutrons form a tiny core called the nucleus, and the electrons form a cloud around the nucleus.
- The number of protons in the nucleus is called the atomic number.
- Each atom has a tiny mass equal to the number of protons plus the number of

neutrons (in atomic mass units, amu).

- A mole is a specific number of atoms (or molecules). A mole is the number of atoms it takes to make the total weight in grams equal the atomic mass (in amu). The number of moles of atoms can be determined by weighing the sample (in grams) and dividing by the atomic mass (in amu).

- All the elements are gathered together in the periodic table of elements. Each element is listed by atomic number. Each row of the periodic table is called a period, and each column is called a group. The periodic table is organized so that the columns are groups of elements with similar properties.

- Among the most important groups are the alkali metals, the halogens, and the noble gases. The periodic table is also divided into two large blocks called the main group and the transition elements.

1.5 STUDY QUESTIONS

1. What are the three basic particles in an atom? Which of these have electric charge, and what charge does each have?

2. An unknown element is found to have 150 protons in its nucleus. How many electrons does it have?

3. An atom is found to have 24 protons and 30 neutrons in its nucleus. What is its atomic mass (in amu)?

4. Identify each of the following elements and give its atomic symbol:

a. Atomic number 17

b. Atomic number 6

c. Atomic number 3

d. Atomic number 7

e. Atomic number 80

5. Which of the following are halogens: atomic number 7, atomic number 12, atomic number 9, atomic number 17, atomic number 35?

6. Which of the following are noble gases: atomic number 10, atomic number 20, atomic number 36, atomic number 2, atomic number 18?

7. Which of the following are alkali metals: atomic number 2, atomic number 3, atomic number 11, atomic number 19, atomic number 21, atomic number 37?

8. How many moles of atoms are in 10.81 grams of boron?

9. How many grams would 3 moles of boron weigh?

10. Ammonia molecules, NH_3, have 1 nitrogen atom and 3 hydrogen atoms. How many grams would 1 mole of ammonia weigh?

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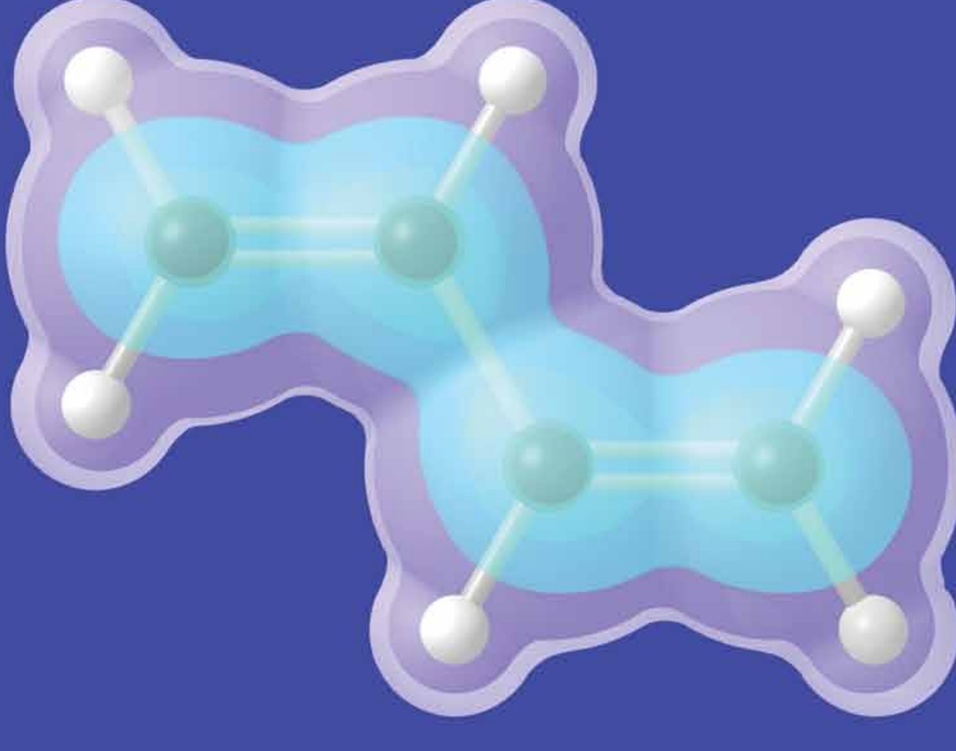

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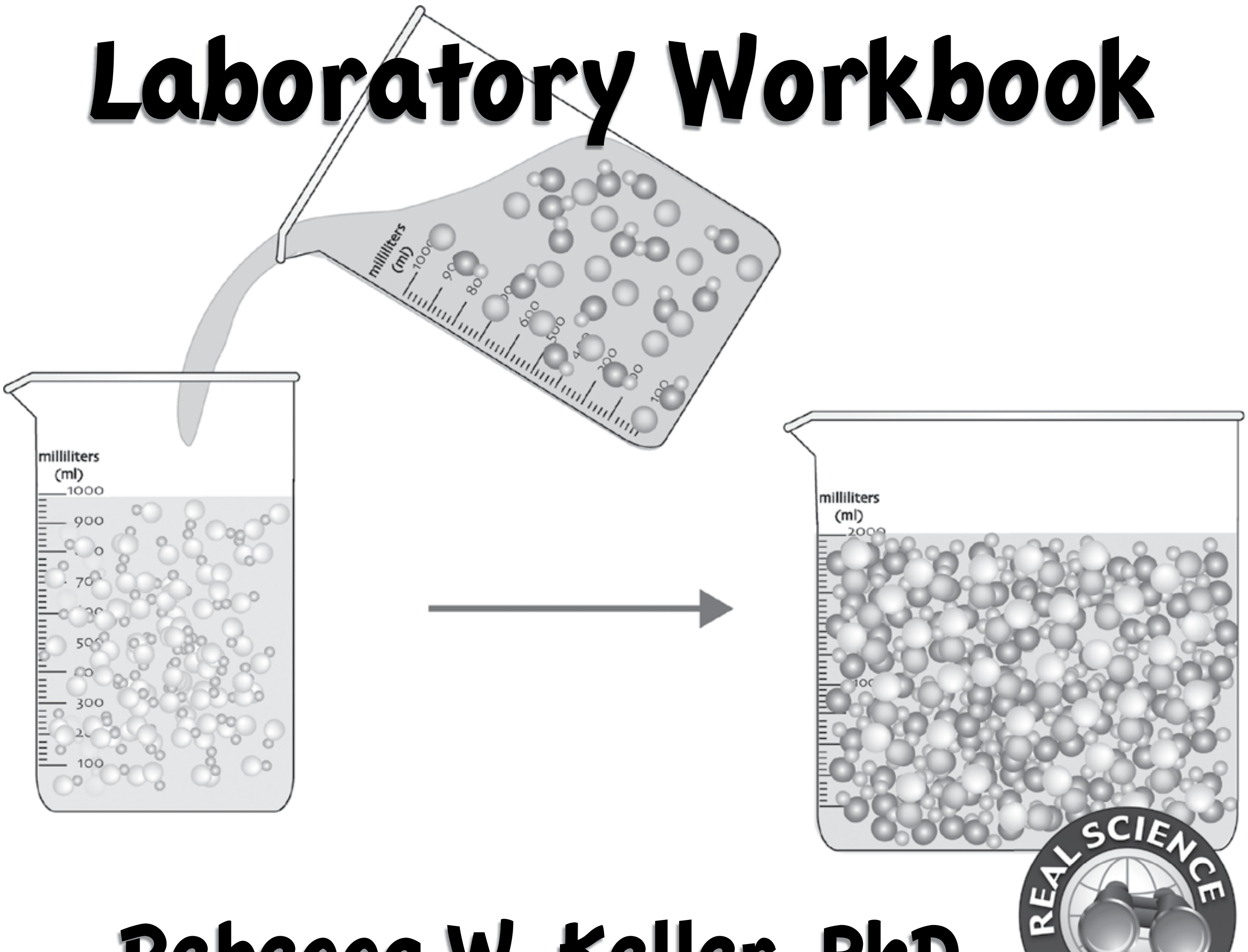
High School Chemistry

Laboratory Workbook



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ON YOUR OWN
A Note From the Author About the Experiments

All of the experiments in the *Focus On High School Chemistry Laboratory Workbook* are essentially "on your own." One of the most essential features of science is the ability to problem solve and create new ideas. You can't do this unless you try. Sometimes it seems like there are not very many chances in life to try new ideas, so I decided to let these experiments be created by you, the student.

At first this is going to be frustrating. Staring at a blank "Experiment" with only a hint or two to work from may seem unfair. You may be nervous about doing it "right," or you may be confused about how to get started. To ease you into the idea, I have written most of the first three experiments for you, but by Experiment 4, you are entirely on your own. The Teacher's Manual contains a "Sample Experiment" for each of the ten chapters if you get really stuck, but resist the temptation to simply follow the teacher's recipe—what do they know?

It may seem odd for a scientist to say this, but there is no one right way to do an experiment. There are better ways to test a hypothesis, just as there are better ways to stick a note to your mom on the corner of her computer screen. (Which is best: duct tape, scotch tape, Elmer's glue, super glue, bubble gum, or Post-Its? Did you know that the glue on Post-Its was glue that "didn't work"?). So yes, there are better ways to do everything, but that is why you have to "experiment" to find out what works best.

Most importantly, have fun. You get to do your own experiments, and in the process I hope you will *discover* real science.

Rebecca W. Keller, PhD

Laboratory Safety

Most of these experiments use household items. However, some items, such as iodine, are extremely poisonous. Extra care should be taken while working with all chemicals in this series of experiments. The following are some general laboratory precautions that should be applied to the home laboratory:

- Never put things in your mouth without explicit instructions to do so. This means that food items should not be eaten unless tasting or eating is part of the experiment.

- Use safety glasses while using glass objects or strong chemicals such as bleach.

- Wash hands before and after handling chemicals.

- Use adult supervision while working with iodine and while conducting any step requiring a stove.

Contents

EXPERIMENT 1: LOW SODIUM

You go to the family doctor, and he decides to put you on a special diet. He tells you that you have been eating too much sodium. He is an old chemist, and he tells you not to eat more than 0.01 mole of sodium per day. This sounds pretty easy until you go home and find out that all of the food items list the amount of sodium in mg (milligrams). How do you follow the doctor's orders? Which foods can you eat?

HINTS:

First determine the atomic weight of sodium. It is on the periodic table and the quantity is given as grams per mole (grams/mole). Record this quantity here: ______________________

Remember that the atomic weight tells you how many grams of an element are in one mole, but you need to find out how many milligrams are in 0.01 mole. To find out how many milligrams of sodium are in 0.01 mole, first convert grams of sodium in one mole to milligrams of sodium (1000 milligram = 1 gram) in one mole and then multiply by 0.01 mole. This will give you milligrams of sodium in 0.01 mole.

Do your calculation here:

milligrams (mg) of sodium in 0.01 mole = ____

Now set up your experiment.

Experiment 1: ______________________________ **Date:** ____________

Objective: ______________________________

Hypothesis: ______________________________

I. List the materials you need.

MATERIALS

____

____

____

____

____

____

II. Write out the steps of your experiment in as much detail as possible.

EXPERIMENT

1. ______________________________

____

____

2. ______________________________

____

____

3. ______________________________

____

____

4. ______________________________

____

____

5. ______________________________

____

____

III. Record your results.

RESULTS

Food Item	Serving size	Sodium (in milligrams)

IV. Discuss your results and write your conclusions.

CONCLUSIONS

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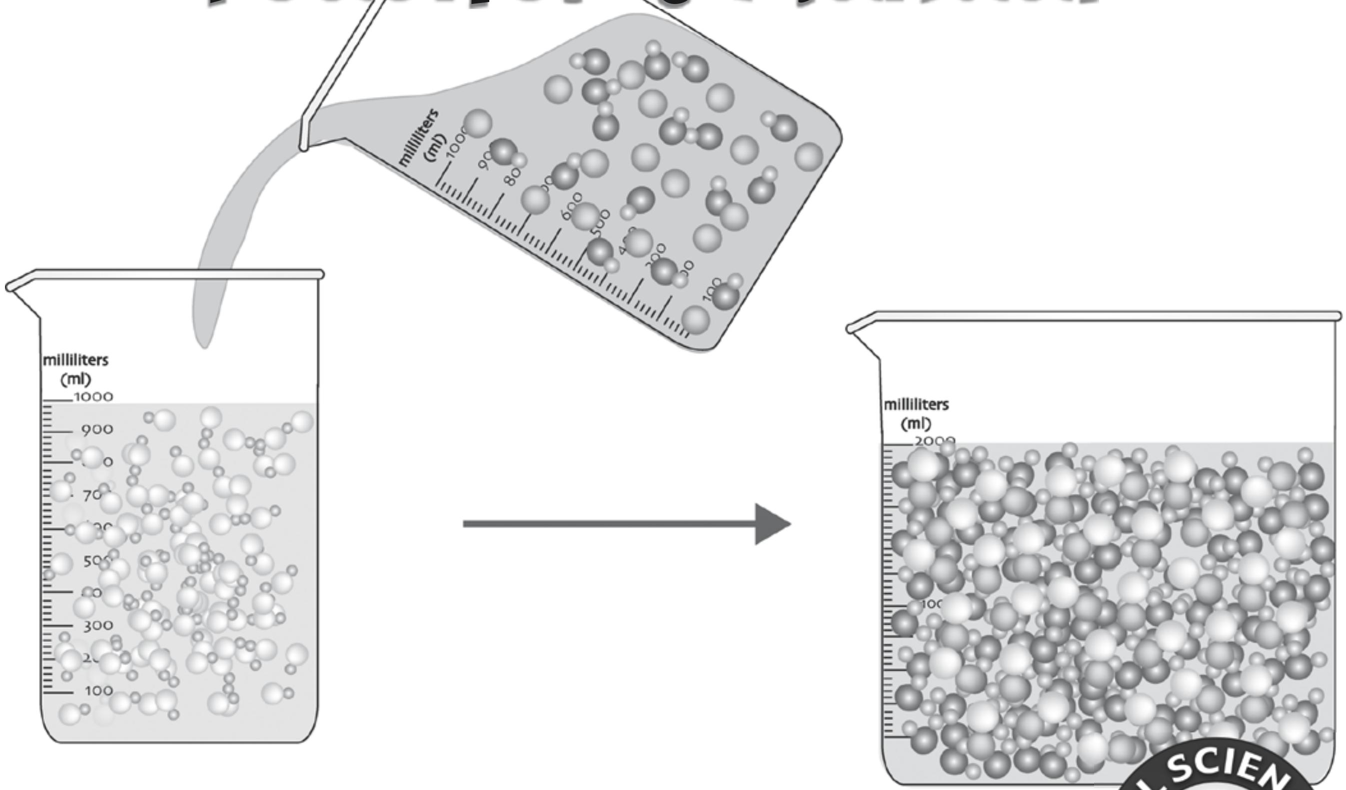

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To the Teacher

In this teacher's manual you will find all of the answers to the Study Questions and Practice Problems in the *Focus On High School Chemistry Student Textbook*. You will also find a "Suggested Experiment" for each chapter.

In the *Focus On High School Chemistry Laboratory Workbook*, I encourage the students to design their own experiments. I want students to learn how to create new scientific ideas, and the only way to get them to do that is to encourage them to try. Anyone can find a number of science "experiments" or recipes to follow in a variety of science lab textbooks and on the internet. But the point of scientific investigation is to discover what is not yet known, and some of the most exciting science is discovered by scientists creating new ways to do new things that no one has tried before.

It is important for teachers to give students the freedom to think for themselves and, within reason, to experiment with ideas that you, the teacher, may know won't work. If your students decide to use bubble gum instead of masking tape to attach a test tube to a jar, let them try. It may work beautifully or it may fail and ruin the whole experiment, but at least they tried and in the end they may know quite a lot more about bubble gum than they would have without trying. It is fine to guide them in order to help them find better ways to conduct scientific experiments, such as the use of controls, but don't get overly concerned that they do the experiment "right." The process of learning through *experimentation* is the most important aspect of these experiments.

Also, as much as possible, answer their questions with more questions. Use lots of "How," "Why," and "What" questions.

Student: "How do I get the test tube to stick to the inside of the jar that contains water?"

Teacher: "How could you do it?"

Student: "I think I'd like to use bubble gum."

Teacher: "Why do you think that will work?"

Student: "Well, because bubble gum is still sticky inside my mouth, and my mouth has water in it, so it should stick in water."

Their answers may surprise you.

Most importantly — be safe and have fun!

Rebecca W. Keller, PhD

Laboratory Safety

Most of these experiments use household items. However, some items, such as iodine, are extremely poisonous. Extra care should be taken while working with all chemicals in this series of experiments. The following are some general laboratory precautions that should be applied to the home laboratory:

- Never put things in your mouth without explicit instructions to do so. This means that food items should not be eaten unless tasting or eating is part of the experiment.
- Use safety glasses while using glass objects or strong chemicals such as bleach.
- Wash hands before and after handling chemicals.
- Use adult supervision while working with iodine and while conducting any step requiring a stove.

Contents

Chapter 1: Matter, Mass, and Moles

Text reading time	2 hours
Experiment time	2-3 hours

Suggested Materials :

any packaged materials that have the salt content listed on the side of the box in milligrams (mg)

Answers to Study Questions

1. protons, neutrons, and electrons. Protons have a positive (+) charge and electrons have a negative (-) charge. Neutrons have no charge.

2. 150

3. 54 amu

4. a. chlorine Cl
   b. carbon C
   c. lithium Li
   d. nitrogen N
   e. mercury Hg

5 atomic number 9, fluorine; atomic number 17, chlorine; atomic number 35, bromine

6. atomic number 10, neon; atomic number 36, krypton; atomic number 2, helium; atomic number 18, argon

7 atomic number 3, lithium; atomic number 11, sodium; atomic number 19, potassium; atomic number 37, rubidium.

8. one mole

9. One mole weighs 10.81 grams, so 3 moles would be: 3 x 10.81 = 32.43 grams.

10. To find how much one mole of ammonia weighs, first look on the periodic table and find the atomic weight for each atom in the ammonia molecule.

    one mole of nitrogen weighs 14.01 grams
    one mole of hydrogen weighs 1.00 gram

    Next, write down how many of each kind of atom ammonia has.

    ammonia has one (1) nitrogen atom and three (3) hydrogen atoms:

    Now, plug these numbers into an equation where the number of each atom is multiplied by its atomic weight.

    One (1) nitrogen atom times (x) 14.01 grams, plus (+) three (3) hydrogen atoms times (x) 1.001 equals (=) 17.01 grams.

    [(1 N) x (14.01 grams)] + [(3 H) x (1.00 gram)] = 14.01 grams + 3 grams = 17.01 grams

    one mole of ammonia weighs 17.01 grams.

Instructions for Experiment 1: Low Sodium

In this chapter students learned about matter, mass, and moles. The experiment outlined below will help students further explore these concepts.

The student will need to learn dimensional analysis to perform this experiment. A full discussion of dimensional analysis is given in Appendix D. Have the student read Appendix D before doing the experiment.

In this experiment, the student is given a hypothetical request by their family doctor to limit their sodium intake. The limit is expressed in moles. They will discover that all of the food products list the sodium amount in milligrams (mg). Help the student think about how to solve this dilemma.

A “hint” is provided to help the student get started.

If your student gets stuck or frustrated, help them think through the experiment by asking the following questions:

1. What can you call your experiment? What are you trying to find out with this experiment?

2. What is an “objective?” Specifically, what is your objective with this experiment? What did the doctor request?

3. What is a “hypothesis?” What foods do you think you may or may not be able to eat?

4. How would you write the steps for the experiment? What do you think you should do first?

5. How could you organize the information from the food labels? Can you put the information in a table or a graph? Which information do you think you should look for? What is the “daily recommended allowance?” What is the serving size for each item?

6. How many food items do you think you should check?

7. What if all of the food items have too much sodium? Should you look for other food items? Do you think you could eat less of each?

8. What is a “conclusion?” Did you prove or disprove your hypothesis? How can you tell?

EXPERIMENT 1: LOW SODIUM

You go to the family doctor, and he decides to put you on a special diet. He tells you that you have been eating too much sodium. He is an old chemist, and he tells you not to eat more than 0.01 mole of sodium per day. This sounds pretty easy until you go home and find out that all of the food items list the amount of sodium in mg (milligrams). How do you follow the doctor's orders? Which foods can you eat?

HINTS:

First determine the atomic weight of sodium. It is on the periodic table and the quantity is given as grams per mole (grams/mole). Record this quantity here: *22.99 grams/mole*

Remember that the atomic weight tells you how many grams of an element are in one mole, but you need to find out how many milligrams are in 0.01 mole. To find out how many milligrams of sodium are in 0.01 mole, first convert grams of sodium in one mole to milligrams of sodium (1000 milligram = 1 gram) in one mole and then multiply by 0.01 mole. This will give you milligrams of sodium in 0.01 mole.

Do your calculation here:

There are 22.99 grams of sodium in one mole.

22.99 grams (in one mole) x 1000 milligrams/gram = 22990 milligrams (in one mole)

$$\boxed{(22.99\ \cancel{\text{grams}}) \times \left(1000\ \frac{\text{milligrams}}{\cancel{\text{gram}}}\right) = 22990\ \text{milligrams}}$$

22990 milligrams/mole x 0.01 mole = 229.9 milligrams

$$\boxed{\left(\frac{22990\ \text{mg}}{\cancel{\text{mole}}}\right) \times (0.01\ \cancel{\text{mole}}) = 229.9\ \text{mg (milligrams)}}$$

milligrams (mg) of sodium in 0.01 mole = *229.9 mg sodium*

Now set up your experiment.

Experiment 1: *What can I eat?* **Date:** ____________

Objective: *To determine which foods contain less than 0.01 mole of sodium per serving*

Hypothesis: *I will be able to eat cereal but not peanut butter.*

I. List the materials you need.

MATERIALS

1. Calculation from page 1 of this experiment

2. Several food item package containers

3. Pen

II. Write out the steps of your experiment in as much detail as possible.

EXPERIMENT

1. *First I will record how much sodium is in 0.01 mole.*
2. *Next, I will make a list of several items and record the amount of sodium in each.*
3. *I will then compare the amount of sodium in each food item with the limit.*
4. *I will determine the food items below the limit and list these as permissible foods.*
5.

III. Record your results.

RESULTS

Food Item	Serving size	Sodium (in milligrams)
Raisin Bran cereal	*1 cup*	*350 mg*
Nature Valley Granola Bars	*2 bars*	*160 mg*
Jiff Peanut Butter	*2 Tbsp*	*150 mg*
Chicken of the Sea tuna	*2 oz.*	*250 mg*
Baked Beans	*1/2 cup*	*550 mg*

Foods that are below 229 mg sodium:

Granola bars - 2 bars, 160 mg sodium

Peanut butter - 2 Tbsp, 150 mg sodium

IV. Discuss your results and write your conclusions.

CONCLUSIONS

If I follow the serving size suggestions for each food item, I will only be allowed to eat the peanut butter and granola bars.

My hypothesis was incorrect. I am not able to eat the cereal at the suggested serving size.

I can reduce the serving size for the food items with high sodium and still be within the 0.01 mole limit.

I can eat less than 1/4 cup of baked beans or 1 oz. of tuna or 1/2 cup of bran cereal.