Chapter 2: Atoms, Elements and the Periodic Table

The Nuclear Model of the Atom

3.2 Elements and Symbols

- Substances which can not be broken down into simpler substances by chemical reactions are called elements.
- The smallest particle size of an element is an atom.
- There are 92 natural elements found on earth, and a number of synthetic elements.
- All elements above 82 are unstable and undergo radioactive decay.

Symbols

Elements are represented by symbols or formulas for simplicity.
- For most elements, the symbol for the element is taken from one or two of the first three letters of the name, e.g. lithium = Li; titanium = Ti; boron = B; nitrogen = N; chromium = Cr
- The first letter (only) is capitalized. Any others must be lower case.
- The symbols for 11 elements do not correspond to their English name because they originally had Latin names, e.g. Na = sodium (from Natrium); K = potassium (from Kalium) You must memorize them.

Suggestion for Learning: Highlight the 11 exceptions on your periodic table

See: http://www.dayah.com/periodic/

Tom Leher had something to say about “The Elements”: How many are there? (count fast) http://www.privatehand.com/flash/elements.html

(Demo: Start “Total” demo)

Formulas

- For most elements, the symbol and chemical formula are the same, e.g. helium = He; barium = Ba; carbon = C.
- Seven elements exist in nature as diatomic molecules rather than individual atoms, e.g. H₂, N₂, O₂, and the halogens F₂, Cl₂, Br₂, I₂. The subscript 2 indicates that two atoms make up the molecule. You must memorize these diatomic elements.
Naming

The name of an element is always a single word, e.g. H = hydrogen, O = oxygen, Ca = calcium.

When elements combine to form compounds, the names are almost always two words, e.g. H₂O = dihydrogen monoxide (also given the common name: water) e.g. CaO = calcium oxide

- **Elements combine in simple whole number ratios to form molecules or compounds.**
  - For a given compound or molecule, these ratios are always precisely the same, giving rise to the **Law of Definite Composition**.
- Compound formulas are written with symbols of the elements involved.
- Subscripts are used to indicate the ratio of atoms that combine to form the compound, e.g. H₂O for water.

**Examples: Names, Symbols, Formulas**

<table>
<thead>
<tr>
<th>Element Present</th>
<th>Elemental State at 25 °C</th>
<th>Molecule</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen</td>
<td>colorless gas</td>
<td>H₂</td>
</tr>
<tr>
<td>nitrogen</td>
<td>colorless gas</td>
<td>N₂</td>
</tr>
<tr>
<td>oxygen</td>
<td>pale blue gas</td>
<td>O₂</td>
</tr>
<tr>
<td>fluorine</td>
<td>pale yellow gas</td>
<td>F₂</td>
</tr>
<tr>
<td>chlorine</td>
<td>pale green gas</td>
<td>Cl₂</td>
</tr>
<tr>
<td>bromine</td>
<td>reddish brown liquid</td>
<td>Br₂</td>
</tr>
<tr>
<td>iodine</td>
<td>lustrous, dark purple solid</td>
<td>I₂</td>
</tr>
<tr>
<td>Substance</td>
<td>Symbol or Formula</td>
<td>Natural Form</td>
</tr>
<tr>
<td>-------------------</td>
<td>-------------------</td>
<td>--------------------</td>
</tr>
<tr>
<td>Helium</td>
<td>He</td>
<td>Atom</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>Atom/Crystalline solid</td>
</tr>
<tr>
<td>Barium</td>
<td>Ba</td>
<td>Atom/Crystalline solid</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H₂</td>
<td>Molecule</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O₂</td>
<td>Molecule</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl₂</td>
<td>Molecule</td>
</tr>
<tr>
<td>Hydrogen chloride</td>
<td>HCl</td>
<td>Molecule</td>
</tr>
</tbody>
</table>
3.3 The Periodic Table: Introduction

Symbolic Representation of Elements: Atomic Number and Atomic Mass (see: http://www.webelements.com/)

Organization of the Periodic Table:

The key to understanding chemistry and chemical reactions is to understand the underlying reasons why the periodic table is organized the way it is.

It has been known for many centuries that certain elements exhibit recurring or periodically repeating properties. Based on those properties, elements were combined into Groups, and the Groups were given names as shown below.

Look up: anatomy_of_the_periodic_table.pdf
Elemental Symbols and the Periodic Table

The primary arrangement of the elements in the Periodic table was developed by a Russian scientist and teacher named Mendeleev trying to come up with a way to organize the elements for his students to understand. The table was initially organized by writing the name of each element on separate slips of paper, like you might prepare study cards. Then, Mendeleev started grouping elements by the similarity in properties, and then by the mass of the elements. The resulting table ultimately developed into one of the major achievements of his time.

http://www.dayah.com/periodic/
The initial organization of the Periodic table was due to the observation that elements can be grouped together based on common properties. Also, when elements are arranged in order of increasing atomic mass, certain patterns of behavior also periodically repeat.

In the end, the correct organizational property of the elements is based on the atomic number (number of protons) rather than the atomic mass (although both agree pretty well), and also the following:
- Elements with similar chemical and physical properties are grouped together so that they fall into the same vertical column.
- **Columns of elements are called Groups**
- Horizontal rows of elements are called **Periods**. A new Period is started when the chemical properties of the element “periodically” repeat and can be placed underneath a previous element. [each Period corresponds to a principal quantum number, or the highest occupied electron shell, discussed in Chapter 11]
- Each new Period of elements shows the same pattern of properties observed in the previous Period
- Position of the element on the table allows us to predict properties of that element

**Group “Family” Names**
Some groups also have a common group name in addition to the group number, as shown by the color coding of the periodic table above. The initial ones to learn are:
- Group IA or 1A: alkali metals
- Group IIA or 2A: alkaline earth metals
- Group VIIA or 7A: halogens
- Group VIIIA or 8A: noble gases

The Roman numeral system is useful in allowing us to predict chemical behavior, as we will see below.
The table as written here has an odd initial appearance. It starts off with a horizontal row of 18 numbered boxes, but only **two elements** are placed in that first row (Period 1) – 16 boxes are empty. Two additional rows (Periods 2 and 3) of **8 elements** each, with 10 boxes empty in each, follow the first Period. Finally, starting with the fourth row (Period 4), all **18** boxes are filled with elements. This repeats again for Period 5.

Initially, Mendeleev’s table appeared to be nothing more than an organizational table. However, it turned out to have an important **physical** significance. Its importance was fully established when it proved capable of predicting the existence and physical properties of elements which had not yet been discovered.
One of the ultimate beauties of this table for us is that it allows us to predict the chemical behavior of an element by examining where in the table it is located. Its horizontal position identifies which vertical Group it belongs to, and elements within a Group share similar chemical properties [elements in the same Group have the same valence electron configuration, discussed in Chapter 11].

Note that the Groups are numbered using two different numbering systems (there are actually three numbering systems). We will focus on the American system using Roman/Arabic numerals and letter designations for reasons that will become clear. The IUPAC system uses numbers only.

Animated Tutorial:
http://zircon.mcli.dist.maricopa.edu/mlx/warehouse/00601-00700/00696/Chemistry.swf
Animated tutorial: periods, groups, solids, liquids, gases, etc.

Video: The Periodic Table (the first 17 minutes only):
Complete the answer sheet that will be provided to you.

Which elements do you need to memorize?

You will need to memorize the name and symbol for the first 18 to begin with followed by the most common, shown below, and then the remainder in Formula – 1 Chemistry 4 Supplement.
Why must I memorize the symbols for the elements?
It may help to consider the similarity of our English language alphabet to the chemical “alphabet”. With 26 letters, we can create millions of words. For example, two letters forming ADD, and DAD, two dramatically different words. Chemistry could be viewed somewhat like learning a foreign language where different symbols are used, e.g. learning Chinese, Arabic, or many other languages requires learning a new alphabet. Like language, with a couple of dozen chemical elements, we can combine them to make millions of chemical compounds….. e.g. hundreds can be made from the combination of the elements C and H alone…..but you have to learn the chemical alphabet first.

For complete Periodic tables, click on:
Additional tables:
• http://pubs.acs.org/cen/80th/elements.html American Chemical Society: detail on each element

Animated Tutorial: Question and answer interactive learning
• http://preparatorychemistry.com/Bishop_complete_electron_configurations_frames.htm

Types of Elements – Three Types
Metals
• about 75% of all the elements are metals
• Properties: lustrous, malleable, ductile, conduct heat and electricity

Nonmetals
• Properties: dull, brittle, insulators

Metalloids (shown in blue in the Table below)
• also known as semi-metals
• Properties: share properties of both metals & nonmetals because they are on the border between metals and nonmetals.

You can readily distinguish between metals and non-metals by seeing which side of the angled dark line or “stair step” the element falls on, as shown
below. To the left of the “stair step”, everything is a metal; to the right is a non-metal.

**Exercise:**

- Write the symbols for the following elements including atomic number and atomic mass.
- Identify which period, group and family (if any) the element belongs to. Specify whether metal, non-metal, or metalloid.

- Strontium
- Silicon
- Xenon

Organize your answers into a table, e.g.

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Atomic Number</th>
<th>Metal, Non-metal, or metalloid?</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Important Elements to Know:**

Handout
- Formulas – 1: Important Elements and Their Symbols. *You must memorize these by the end of Chapter 2.* You should start to learn these now, beginning with the first 18.
- Periodic Table. For more detail, see: [http://www.dayah.com/periodic/](http://www.dayah.com/periodic/)

**Essential elements:** Thirty-three of the elements are essential for life. The names of the elements are contracted into **symbols** for simplicity.
Questions:

- Which element do you think is the most abundant element in the human body?

- Which one do you think is the most abundant element in the earth’s crust?

*Refer to the table inside the front cover of the textbook for the names corresponding to the symbol for the element.*

*Note: you will not have a periodic table with names for quizzes and exams.*
Relative Abundance of Elements in the Human Body

Abundance is determined as the percentage found in nature

1. Percent composition by number of atoms:

<table>
<thead>
<tr>
<th>Element</th>
<th>Percentage Composition by Number of Atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>63.0</td>
</tr>
<tr>
<td>Oxygen</td>
<td>25.5</td>
</tr>
<tr>
<td>Carbon</td>
<td>9.45</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>1.35</td>
</tr>
<tr>
<td>Calcium</td>
<td>0.31</td>
</tr>
<tr>
<td>Phosphorus</td>
<td>0.22</td>
</tr>
<tr>
<td>Chlorine</td>
<td>0.057</td>
</tr>
<tr>
<td>Sulfur</td>
<td>0.049</td>
</tr>
<tr>
<td>Sodium</td>
<td>0.041</td>
</tr>
<tr>
<td>Potassium</td>
<td>0.026</td>
</tr>
<tr>
<td>Magnesium</td>
<td>0.013</td>
</tr>
</tbody>
</table>

These four elements make up 99.3% of the atoms in your body.

2. Percent composition by the mass of the atoms:
Percent composition of elements on earth (by mass):

<table>
<thead>
<tr>
<th>Major Elements</th>
<th>Mass Percent</th>
<th>Trace Elements (in alphabetical order)</th>
</tr>
</thead>
<tbody>
<tr>
<td>oxygen</td>
<td>65.0</td>
<td>arsenic</td>
</tr>
<tr>
<td>carbon</td>
<td>18.0</td>
<td>chromium</td>
</tr>
<tr>
<td>hydrogen</td>
<td>10.0</td>
<td>cobalt</td>
</tr>
<tr>
<td>nitrogen</td>
<td>3.0</td>
<td>copper</td>
</tr>
<tr>
<td>calcium</td>
<td>1.4</td>
<td>fluorine</td>
</tr>
<tr>
<td>phosphorus</td>
<td>1.0</td>
<td>iodine</td>
</tr>
<tr>
<td>magnesium</td>
<td>0.50</td>
<td>manganese</td>
</tr>
<tr>
<td>potassium</td>
<td>0.34</td>
<td>molybdenum</td>
</tr>
<tr>
<td>sulfur</td>
<td>0.26</td>
<td>nickel</td>
</tr>
<tr>
<td>sodium</td>
<td>0.14</td>
<td>selenium</td>
</tr>
<tr>
<td>chlorine</td>
<td>0.14</td>
<td>silicon</td>
</tr>
<tr>
<td>iron</td>
<td>0.004</td>
<td>vanadium</td>
</tr>
<tr>
<td>zinc</td>
<td>0.003</td>
<td></td>
</tr>
</tbody>
</table>
– oxygen most abundant element (by mass) on earth and also in the human body
– the abundance and form of an element varies in different parts of the environment

Alphabetical list of the most common elements:
Note: A more complete list is on the front cover page of your textbook.

Very few elements exist on their own, e.g. Au, some Cu, some S, O₂, and N₂ to name a few. Most substances are chemical combinations of elements. These are called compounds.

- Compounds are made of elements
- Compounds can be broken down into elements
- Properties of the compound are not related to the properties of the elements that compose it
- Same chemical composition at all times

Exercise: Elements compounds mixtures
3.4 The Atom

Dalton’s Atomic Theory

**The Greek Atomos**
Around 440 BC, Leucippus of Miletus originated the concept of the atom. He and his pupil, Democritus (c460-371 BC) of Abdera, in ancient Greece, refined and extended it in future years. This concept was not based on any data or experimentation. This was a philosophical argument based on the logical reasoning that if you start to divide or split something, there must be a stopping point beyond which you cannot divide something further. Democritus quotes Leucippus: "The atomists hold that splitting stops when it reaches indivisible particles and does not go on infinitely." In other words, there is a lower limit to the division of matter beyond which we cannot go. *Our word atom therefore comes from atomos, a Greek word meaning uncuttable or unsplittable.*

It was almost 2000 years later before this concept of the atom acquired an experimental basis culminating in John Dalton’s atomic theory.

**John Dalton (1768 – 1828)**
Dalton, an English schoolteacher, studied the mass ratios of the elements involved in numerous chemical reactions. He combined his own observations with those of Lavoisier (Law of Conservation of Mass), Proust (Law of Definite Proportions), and the idea that elements cannot be broken down into simpler substances to formulate the beginning of modern atomic theory. In 1808 he proposed:

**Dalton’s Atomic Theory: The Beginning of a Scientific Revolution**
- Elements are composed of tiny indivisible particles called atoms
- All atoms of a given element are identical - they all have the same size, mass, and chemical properties
- Atoms of one element are different from those of any other element
- Atoms are indestructible and retain their identity in chemical reactions. - all atoms present at beginning are present at the end - atoms are not created or destroyed, just rearranged - atoms of one element cannot change into atoms of another element
- **Compounds** are formed by combining different atoms in specific whole number ratios.
Dalton envisioned compounds as collections of atoms such as the NO, NO$_2$ and N$_2$O, shown here:

See also: DaltonAtomicTheoryOverhead.pdf

*Dalton’s atomic theory* provided a proposed explanation for the laws such as the law of constant composition (Proust). However, *a theory is not proof* [of the existence of atoms.] The proof of the existence of atoms did not occur until the early 1900’s, when an obscure young Swiss patent clerk examined some 80 year-old data from a botanist by the name of Brown. His work examining “Brownian motion” data (and also the photoelectric effect, chapter 11), not his theory of relativity, led to the Nobel Prize in Physics for Albert Einstein.

Note also that Dalton’s theory did not explain at all why or how atoms stick together, or why molecules stick together. *The quest for the atomic force,* discussed beginning in Chapter 4, continues still today.

**Coulomb’s Law**
Charles Coulomb, a contemporary of Dalton, made an important discovery about charged particles which will become important in the later exploration of the structure of the atom. He found that two identically charged objects repel each other, whereas oppositely charged object are attracted each other. The force of this interaction is inversely proportional to the square of the distance (r) separating them, i.e.

\[ F \propto \frac{1}{r^2} \]

Like charges repel:

\[ - \leftrightarrow - \]

Opposite charges attract:

\[ + \rightarrow - \]

Stronger charges attract more strongly:

\[ ++ \leftrightarrow -- \]

Coulomb’s law is closely analogous to Newton’s law of gravity, where the force of gravity is proportional to the inverse cube of the distance.

The Structure of the Atom
So, is the atom indivisible? To learn about the structure of the atom is like reading a mystery novel. It’s a story about an avalanche of discoveries which have enormous implications for understanding ourselves and the universe we live in. It’s an ongoing story whose ending is yet to be written. The scientific process of discovery is outlined below.

5.2 Subatomic Particles

Cathode Rays: Discovery of the Electron
In the mid-1800’s, an unrelated engineering development lead to several decades of study of an odd phenomenon.

A study of this phenomenon, which for decades was used as a parlor trick, lead to experiments culminating in the work of J.J. Thompson in 1897 which showed that the “indivisible” atom was indeed divisible.

That engineering accomplishment was the mercury vacuum pump, a pump that could remove most air from containers. When this pump removed most of the air from a special glass tube containing metal electrodes at each end, and the electrodes were “charged” by attaching them to a battery, a mysterious glow was observed between the electrodes in the tube. The glow became known as a cathode ray since it originated from the cathode or negatively-charged electrode, and the tube became known as Crookes tube or later a cathode ray tube (CRT). This glow was a puzzling phenomenon. Was it light? Was it matter?
J.J. Thompson (1856 – 1940) applied the scientific process to a study of cathode rays, and his results are summarized below.

**To summarize,**
- All atoms contain inside of them a tiny, negatively charged particle.
- This particle, called an electron, is identical in all atoms.
- The electron is a basic building block of all atoms.
- It must be possible to subdivide atoms. One of the assumptions of Dalton’s model had been destroyed. Atoms could indeed be divided further.

In addition, Thompson showed that this particle was at least **1000 times smaller** than the smallest known atom or ion (the Hydrogen ion).

**Where does it stop?** We are still trying to answer that question today.

**Also, it eventually raised some challenging new questions:**
- How can atoms be different if they all contain the same parts?
- How can these negative particles coexist in the atom? Wouldn’t they immediately repel each other (Coulomb’s law)?

**The Proton**
If negatively-charged particles were contained within all atoms, and yet all atoms are electrically neutral (that is, atoms do not have a negative charge) it was reasoned that there must be a counterbalancing positive charge present. It remained to Ernest Rutherford to eventually conclusively prove this hypothesis.

This raises another astounding scientific dilemma: how can positive and negative charges coexist in close proximity without annihilating themselves?

**Subatomic Structure: The Planetary Model of the Atom**

**Discovery of the Nuclear Atom:** Rutherford’s Gold Foil Experiment
The discovery of the electron (a negative charge) and a positive charge (proton) within the atom begged the question:

- **How must these particles and their positive and negative charges be arranged within the atom?**
- **What is the internal substructure of an atom?**

Initial speculation was that electrons were distributed like “plums in a pudding” of positive charge, or raisins in a raisin cookie of positive charge.

**The plum pudding or raisin cookie model:** electrons are embedded in a spherical cloud of positive charge.

4.3:
One of the early models of the atom was the plum pudding model.

![Spherical cloud of positive charge](image)

Electrons

However, a brilliant experiment by Ernest Rutherford showed that when data talks, speculation walks. The results were so astounding that most scientists thought Rutherford was playing a practical joke on them when he reported the data.

**Ernest Rutherford (1871 – 1937),** whom Einstein called **“the second Newton”**, was studying the nature of radioactivity – spontaneous emission of radiation - which he determined to be three types of high energy particles. He employed one form of radioactive emission, the alpha (α) particle
emitted by the radioactive element **radon**, to probe the interior structure of the atom.

Alpha particles are positively charged particles. The alpha particles were collimated in a lead box employing a long narrow opening. Traveling at $1.5 \times 10^7$ meters/second (about $1/10^{th}$ the speed of light), the alpha particles were aimed at an extremely thin gold foil (about a 1000 atoms thick), and an apparatus which could detect alpha particles was arrayed around the gold foil, as shown in the figure above. This experiment was thought to be rather mundane, and was given to a new a graduate student, Ernest Marsden.

What were they expecting?
1. The alpha particle (+ charge), if it approached a proton (+ charge) close enough, might be slightly deflected by the repulsion force. The degree of deflection would depend on how close the alpha particle came to the proton.
2. The frequency, distribution, and degree of deflection by the protons would reveal something about the arrangement of particles inside the atom. Presumably, they thought they would simply be confirming the plum pudding/raisin cookie model.
Note: electrons would have no effect because they were so tiny vs. the alpha particle. They also were of opposite charge, so would be attracted to each other rather than be repelled.

**Animation:**
Tutorial/animation: Rutherford experiment

**Gold Foil Results:**

The vast majority of the alpha particles passed right through the gold foil with very little or no deflection, blasting right through the haze of positive charge, as the plum pudding model might predict. About 2% showed varying amounts of deflection indicating they had came close to a proton. However, to the utter amazement of everyone, a few alpha particles were greatly deflected, and about 1/20000 even bounced almost straight back.
To summarize:

- Over 98% of the $\alpha$ particles went straight through.
- About 2% of the $\alpha$ particles went through but were deflected by large angles.
- About 0.01% of the $\alpha$ particles bounced backwards off the gold foil.

What intense electrical forces could have caused an alpha particle traveling at such astronomical speeds to change directions so drastically? And, while also conserving momentum and energy?!

The Plum pudding / Raisin cookie model of the atom could not explain these results.

These results could only be explained by the presence of an incredibly dense, tiny region of positive charge inside the atom, which Rutherford
called the **nucleus of the atom.** He proposed that this nucleus must contain the positive charge in the form of a particle which he called a **proton.** Because atoms have no net charge, the positive charge on a proton must be the same magnitude as the size of the negative charge on the electron.

**Summary: Rutherford’s Atom**
- The atom contains a tiny dense center called the **nucleus**
  - The volume of the nucleus is about 1/10 trillionth the total volume of the atom. In other words, the atom is almost entirely empty space
- The nucleus essentially contains the entire mass of the atom
- The nucleus is positively charged, i.e. it contains all the protons
- The amount of positive charge of the nucleus exactly balances the negative charge of the electrons, i.e. the # protons = # electrons
- The electrons move around in the empty space of the atom surrounding the nucleus

Rutherford’s Atom

Rutherford’s atom started a revolution in our understanding of nature, or perhaps more accurately, revealed how little we knew.
**Exercise:** Rutherford’s atom still leaves us with two major dilemmas. Can you identify them?

So, how **many** units of positive charge reside in the nucleus of each atom?

Using gamma rays (Mosley) and alpha particles (Chadwick and others) in a technique similar to Rutherford, the number of protons in the nucleus of the elements was determined; in this example below, by examining the amount of deflection experienced by the alpha particles as they were fired at different metal foils.

Consider the experiment: logically, smaller atoms would require the alpha particles to come closer to the nucleus to achieve the same amount of deflection by $+/-$ repulsive force compared to larger nuclei with more protons ($+$ charge) as shown in the figure below. So deflection of the particles could be used as a tool to probe the interior of the nucleus.
Experimentally, a fixed angle of alpha particle deflection was chosen, and the ratio of the frequency of alpha particles deflected to that angle was compared for different elements. By this means, the number of protons, and therefore the atomic number, was assigned to most of the elements in the periodic table.

The identity of a given element is defined by its atomic number (Z), which is the number of protons in the nucleus of the atom. In the Periodic table (discussed below), the known elements are arranged in order of increasing Z or atomic number.
To summarize:

- The # of protons determine the element
- The Atomic Number is the # of protons
- The # protons = # electrons (in an uncharged atom)

Elements are represented in the periodic table in the following manner:

<table>
<thead>
<tr>
<th>Atomic #</th>
<th>2</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td></td>
</tr>
<tr>
<td>Atomic mass</td>
<td>4.0026</td>
</tr>
</tbody>
</table>

Examples:

- How many protons in Potassium?
- How many electrons in Potassium?

- How many protons in Au?
- How many electrons in Au?

Modern Atomic Structure

The Neutron:

It was subsequently found that the proton and electron only accounted for about half the mass of the atom. A third, electrically neutral, nuclear particle was therefore proposed and eventually found and named the Neutron. *(The neutron is believed to play an important role in stabilizing the nucleus of atoms.)* Together, these three particles account for essentially all the mass of the atom. The relative mass and the charge of all three subatomic particles are summarized in Table 4.4
The actual mass of a proton or neutron is unimaginably small: approximately $1.67 \times 10^{-24}$ g (simplified to 1 atomic mass unit or 1 amu, discussed below). The actual values are shown below.

### Properties of Three Fundamental Particles

<table>
<thead>
<tr>
<th>Particle / Symbol</th>
<th>Electric Charge</th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Coulombs</td>
<td>Atomic</td>
</tr>
<tr>
<td>Proton $p^+$</td>
<td>$+1.602 \times 10^{-19}$</td>
<td>$+1$</td>
</tr>
<tr>
<td>Neutron $n$</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Electron $e^-$</td>
<td>$-1.602 \times 10^{-19}$</td>
<td>$-1$</td>
</tr>
</tbody>
</table>

We now know the nucleus of an atom is incredibly tiny ($\sim 10^{-13}$ cm) as shown in the figure below. The electrons move about outside the nucleus with an average distance of about $10^{-8}$ cm. Therefore the radius of the atom is about $10^5$ times larger than the radius of the nucleus.
To put this into perspective, if the nucleus of the atom were the size of a marble, and it was placed on the 50 yard line of the Rose Bowl, the diameter of the rest of the atom would cover the entire Rose Bowl. In other words, the vast majority of the atom is empty space where the electrons roam.

For a tutorial:  http://web.jjay.cuny.edu/~acarpi/NSC/3-atoms.htm

For even greater depth on subatomic structure, go to:  
http://particleadventure.org/

So, the symbols for elements in the periodic table show the atomic number (the number of protons) on top. Underneath the symbol is the mass of the element.

<table>
<thead>
<tr>
<th>Atomic #</th>
<th>→</th>
<th>2</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td></td>
<td>4.0026</td>
</tr>
</tbody>
</table>
The total mass of an atom is essentially the sum of the mass of the protons + neutrons. The mass of a proton and a neutron are essentially identical. Recall from above that the proton was given an arbitrary unit of 1 amu (atomic mass unit). Since the number underneath the symbol of the element is given as the atomic mass, this suggests that in the example of He above, there are a total of 2 protons + 2 electrons in each atom. For most elements, the atomic mass number very closely reflects this.

Look at your periodic table for some examples: e.g. determine how many neutrons there are in:
- Oxygen
- Neon
- Sodium
- Sulfur

Why is the atomic mass not a whole number? Also, why do some atoms appear to not follow the examples shown above?

There are two reasons. The first and by far the most important is the existence of isotopes.

**Isotopes**

We know that every atom of a given element has the same number of protons. The number of protons is called the atomic number. It is often represented by the letter \( Z \). The atomic number, then, defines the element, and is the number above the symbol in the Periodic table. Because atoms are electrically neutral, the number of electrons must be the same as the number of protons.

A sample of a pure element can none-the-less have atoms of different masses. This is due to variations in the number of neutrons.
- Atoms which have the same number of protons but different numbers of neutrons are called isotopes.
- Neutrons do not influence the chemical properties of the atom, only the mass of the atom, so variation in the number of neutrons do not alter the element.
• Neutrons are actually part of the nuclear glue that holds the protons together in that tiny space called the nucleus, and prevents them from flying apart. Isotopes, then, are simply variations in the amount of nuclear glue.

To summarize isotope properties:
• All atoms of an element have the same number of protons [and electrons]
  - The number of protons in an atom of a given element is represented by the atomic number found above the symbol for the element in the Periodic Table
• Atoms of an element with different numbers of neutrons are called isotopes
• Because mass number = protons + neutrons, isotopes of an element have different masses
• Isotopes are identified by their mass numbers
• All isotopes of an element are chemically identical and therefore undergo the exact same chemical reactions

Figure 4.10: Two isotopes of sodium.
• The vast majority of elements have isotopes.
• The elements found in nature are mixtures of isotopes.
• The mass of an natural element is a weighted average of the isotopes.

Symbolic Representation of Isotopes: Atomic Number and Mass Number
Note: The convention for writing the symbol for an isotope is quite different from the way the elements are written in the periodic table as we will see below.

The nuclear symbol for the isotope of an element consists of three parts:

- The symbol for the element
- The subscript number is called the atomic number = number of protons.
- The superscript number is called the mass number = protons + neutrons
- The mass number and atomic number are always whole numbers

Note:
- **The number of protons defines the element.** Thus, all carbon atoms have 6 protons. All nitrogen atoms have 7 protons.
  
  \[
  \begin{align*}
  ^{14}_{6}C & \quad ^{15}_{7}N \\
  \end{align*}
  \]

- **Atoms of the same element can, however, have different masses.** Each unique form (isotope) of a given element has a specific number of protons and neutrons. For example:

  \[
  \begin{align*}
  ^{14}_{6}C & \quad ^{13}_{6}C \quad ^{12}_{6}C \\
  \end{align*}
  \]

  - All of these atoms are carbon atoms as shown by the letter symbol “C” and the atomic number of 6
  - Each form of carbon above would be called an isotope of carbon. Sometimes they are referred to as Carbon-14 or Carbon-12.
  - All isotopes of carbon have the same number of protons, but their number of neutrons can differ, hence their mass differs.
- Isotopes are chemically identical; they all enter the same chemical reactions

Example:
- An isotope of Chlorine has a mass number of 37 and an atomic number of 17. How many neutrons does it have? How many electrons?
- If another isotope of Chlorine has a mass number of 35, how many neutrons does it have? How many electrons?

Example:

\[
\begin{array}{ccc}
13 & \text{Ba} \\
5 & \text{N}
\end{array}
\]

How many neutrons are present in this isotope?

Exercise:
Create and complete the following table for isotopes:

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium</td>
<td></td>
<td>11</td>
<td>23</td>
<td></td>
</tr>
<tr>
<td>Nitrogen</td>
<td>(^{15}\text{N})</td>
<td>7</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Lithium</td>
<td>(^{136}\text{Ba})</td>
<td>56</td>
<td></td>
<td>6</td>
</tr>
<tr>
<td>Boron</td>
<td></td>
<td>5</td>
<td>11</td>
<td></td>
</tr>
</tbody>
</table>
The Relative Abundance of Isotopes:
The existence of isotopes, and the relative abundance of each, is measured by a technique known as mass spectrometry. In this apparatus, an element is:

- Vaporized and ionized (an e- is removed, leaving the element with a positively charged ion.
- Accelerated in an electric field to form a beam of ions
- Bent by a strong magnetic field
- Detected by a collection device
- Heavier ions are deflected slightly less than lighter ions.

Instructor accessible video: Videos / 2.2 Mass spectrometry: Determining Atomic Masses [http://www.thinkwell.com/marketing/animations.cfm](http://www.thinkwell.com/marketing/animations.cfm)
Neon is shown to have three isotopes. The most abundant isotope is neon-20 at 90.5% of the total. Since there are two isotopes heavier than a mass of 20, the mass of neon averages out to be 20.18.

Most elements have isotopic forms. Some examples are shown in the table below.
Atomic Mass

The Atomic Mass Unit (amu):

The mass of the proton, neutron, and electron are actually known to a high degree of accuracy:

<table>
<thead>
<tr>
<th>Particle</th>
<th>Mass, g</th>
<th>Mass, amu</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>$1.673 	imes 10^{-24}$</td>
<td>1.007276</td>
</tr>
<tr>
<td>Neutron</td>
<td>$1.675 	imes 10^{-24}$</td>
<td>1.008665</td>
</tr>
<tr>
<td>Electron</td>
<td>$0.0009 	imes 10^{-24}$</td>
<td>0.0005486</td>
</tr>
</tbody>
</table>

However, the mass of an atom cannot be calculated simply by adding up the actual mass of number of its fundamental particles, the protons, neutrons, and electrons. The actual atomic mass is smaller that the sum of its parts. This difference, albeit extremely small, is highly important in understanding atoms. The mass difference is due to the nuclear binding energy required to overcome the repellent forces of the positive charges (proton) which are forced to come into close proximity in forming the nucleus.

Where does this energy come from?
Consider Einstein’s famous equation relating mass and energy:

\[ E = mc^2 \]

Where \( E \) = energy; \( m \) = mass; \( c \) = a constant, the speed of light

Simply put, in forming an atom, some of the particle mass is converted to nuclear energy such as the strong nuclear force, the energy required to form a stable nucleus, and hold everything together. This is the same nuclear energy which is released in nuclear reactions such as nuclear fission when this process is reversed and the nuclei of atoms are broken apart.

Note that this nuclear binding energy mass is not a fixed fraction of any atom’s total mass; rather the nuclear binding energy is unique to each element, and the amount of particle mass converted to energy is also unique. Hence, the mass of each element is not a simple multiple compared to each other element. The mass for each element must be experimentally determined.

The carbon-12 isotope was arbitrarily chosen as the reference element, and assigned a mass of 12 atomic mass units, abbreviated amu. All other elements have their atomic masses experimentally determined by comparison to carbon-12 using the technique called mass spectrometry discussed above.

The atomic mass unit (amu) is arbitrarily defined as one 12\( ^{th} \) the mass of a \( ^{12} \text{C} \) atom. Hence, by definition:

- 12 amu = the mass of a single atom of \( ^{12} \text{C} \)
- 12 amu = known as the atomic mass of \( ^{12} \text{C} \)
- 1 amu = 1/12\( ^{th} \) the mass of a \( ^{12} \text{C} \) atom
- 1 amu = \( 1.992647 \times 10^{-23} \) g / 12
  \[ = 1.6605391 \times 10^{-24} \text{ g} \]

### 3.6 Isotopes and Atomic Mass

When elements are arranged into a chart known as the Periodic Table, the symbols for each element in the periodic table are written differently from the symbols for isotopes for reasons described below. For example:
In the periodic table, each element symbol has the atomic number written above the symbol. A second number, known as the atomic mass, is also written below the symbol. This is because the periodic table presents the elements as they are normally encountered in nature, and the atomic mass (not the mass number) written below the element symbol incorporates the occurrence of isotopes in elements.

11
Na Sodium: 11 = the atomic number; 22.99 = the atomic mass in amu
22.99

6
C Carbon: 6 = the atomic number; 12.01 = the atomic mass in amu
12.01

The mass number discussed for a given isotope is a fixed number for that particular isotope, because it is the sum of the number of protons + neutrons. The atomic mass however is the weighted average of the isotopes which make up a particular element. To calculate the (average) atomic mass for an element, you must know the atomic mass for each isotope, and the frequency of occurrence of each isotope.

For example ~99% of carbon is Carbon-12 and ~1% is Carbon-14. In any sample of carbon on earth you will generally find this ratio of the two isotopes. That is, a sample of carbon it contains both isotopes, so you need a
way of determining the average atomic mass. Calculate the average atomic mass based on the abundances of the isotopes listed above. It should match the atomic mass listed on the periodic table - 12.01.

**Example:**

For the element Chlorine:

![Diagram of Chlorine atom]

The atomic mass is the weighted average of the two isotopes.

**Example:**

The element boron has two isotopes: boron-10 and boron-11.
Boron-10 Isotope

\[
\begin{align*}
\text{Mass} & \rightarrow 10 \\
\text{Atomic} & \rightarrow 5 \text{B}
\end{align*}
\]

Boron-11 Isotope

\[
\begin{align*}
\text{Mass} & \rightarrow 11 \\
\text{Atomic} & \rightarrow 5 \text{B}
\end{align*}
\]

Elemental Boron in Periodic Table

\[
\begin{align*}
\text{Atomic} & \rightarrow 5 \\
\text{B} & \\
\text{Atomic mass} & \rightarrow 10.81
\end{align*}
\]

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Frequency</th>
<th>Mass</th>
<th>Mass Contribution</th>
</tr>
</thead>
<tbody>
<tr>
<td>Boron-10</td>
<td>19.9 %</td>
<td>10.01294 amu</td>
<td>1.99 amu</td>
</tr>
<tr>
<td>Boron-11</td>
<td>80.1 %</td>
<td>11.00931 amu</td>
<td>8.82 amu</td>
</tr>
<tr>
<td>Total</td>
<td>100.0 %</td>
<td>-</td>
<td>10.81 amu</td>
</tr>
</tbody>
</table>

**Problem:**
An element has two isotopes as follows:

<table>
<thead>
<tr>
<th>Atomic Mass (amu)</th>
<th>Percent Abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>62.9298</td>
<td>69.09 %</td>
</tr>
<tr>
<td>64.9278</td>
<td>30.91 %</td>
</tr>
</tbody>
</table>

What element is this?