Introduction

To get a feel for how chemicals react, we must first discuss how these chemicals are put together. In general, we need a sense of how **matter** is put together. Our picture of the structure of matter has been revised over the centuries as we've started to understand the structure and look at smaller and smaller fragments of matter. One of the first useful theories explaining how matter might be put together was Dalton's atomic theory - which dates to the early 1800s. This note pack will discuss Dalton's atomic theory, both in its original form and with modern revisions. You'll also be exposed to the subatomic particles: protons, neutrons, and electrons, and their role in determining an atom's behavior.

Dalton's Atomic Theory

Dalton's ideas date back to a thought experiment (proposed by early philosophers) which asked a few simple questions. What would happen if you took a piece of matter such as a lump of gold and divided it in half? Then, what if you take that half and divide it again. - and again - and again? Would you reach a point where you couldn't divide the gold anymore?

Dalton (and the early philosophers proposing the thought experiment) said yes - at some point, you'd hit particles you couldn't subdivide. These were called **atoms**. The concept of the atom is central to Dalton's atomic theory.

Dalton's theory is conveniently stated as a list of **postulates**, or assertions.

1. All matter is composed of indivisible particles called **atoms**.
2. **Elements** are kinds of matter that are composed of only one kind of atom. The atoms that make up an element have identical properties.
3. **Compounds** are composed of atoms of two or more elements combined in fixed proportions.
4. **Chemical reactions** are rearrangements of atoms that make new compounds. However, the atoms themselves are not created or destroyed.

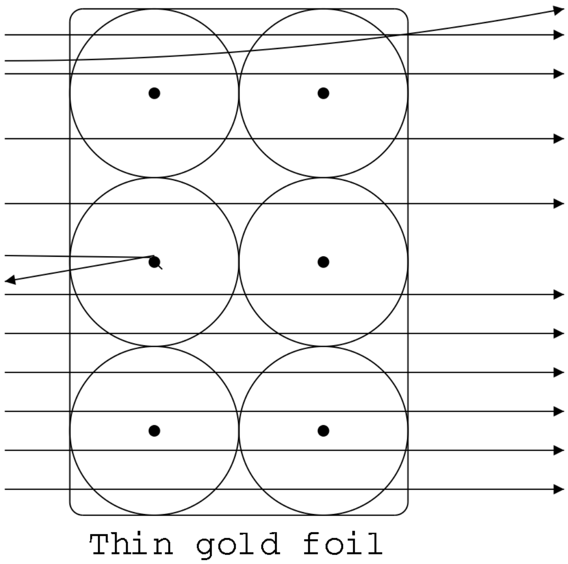
Dalton's theory has held up quite well considering its age. However, we have made some discoveries since then that have required us to modify Dalton's original ideas.

Modification of Dalton's Theory

In the early 1900s, we began to consider that atoms themselves have some sort of structure. What's the difference between a gold atom and a lead atom? Might there be some structure we can determine that's smaller than the atomic level?

Important experiments regarding the structure of the atom was performed by **Rutherford**'s lab. These experiments involved bombarding a thin sheet of gold foil with a beam of fast-moving positively-charged particles. (These particles were alpha particles

- the innards of a helium atom.)

	• Experimentally, more than 99.98% of the particles passed right through the foil to the other side! • Some particles were deflected to a small extent, but not bounced back. • **However, a few of the particles were bounced back, as if they had hit a large, concentrated charge.**
Illustration 1 - Diagram of Rutherford's gold foil experiment	

The results of Rutherford's experiment came as a shock to Rutherford himself, as it showed that both a large amount of the atom's charge and mass were located in an extremely small part of the atom. The atom itself was mostly empty space! From this and other experiments, Rutherford deduced a picture of the atom which we called the **nuclear model**. Atoms had most of their mass (greater than 99% of it) located in a small, dense **nucleus** at the center of the atom. This nucleus carried a positive electrical charge. Surrounding this nucleus was an **electron cloud** of small (very low mass) **electrons**, each carrying a negative charge. The electron cloud contained much of the volume of the atom.

Further experiments led to a more modern notion of an atom containing various subatomic particles: **protons** and **neutrons**. The **proton** was a small but relatively massive positively charged particle found in the nucleus. The **neutron** was similar in mass to the proton, but was electrically neutral.

Particle	*Mass compared to other subatomic particles*	*Description*
proton		positively (+) charged particle located inside atomic nucleus
neutron	A little higher than the proton's mass, but practically equal.	uncharged particle located inside nucleus
electron	tiny compared to protons and neutrons	negatively (-) charged particle located outside of nucleus

The atom, then, might look something like shown below.

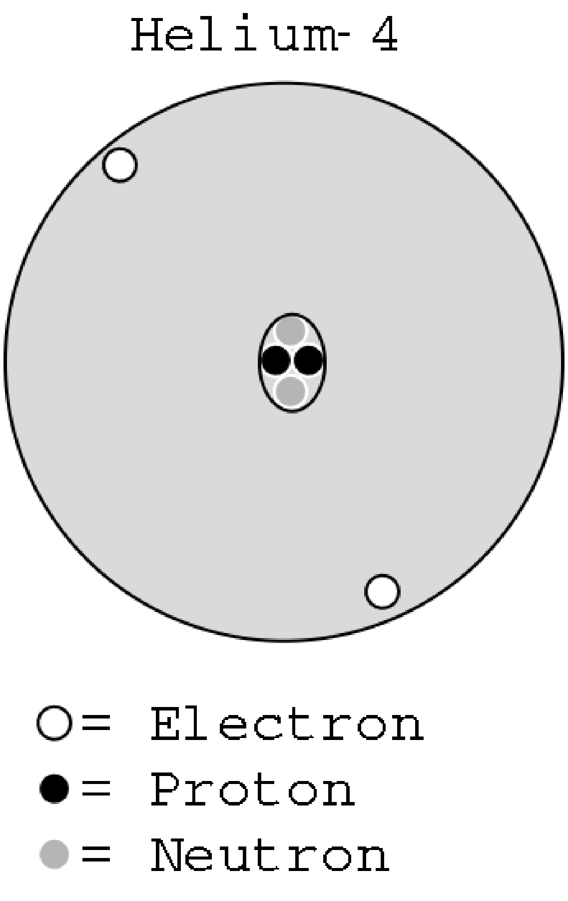


Illustration 2 - Simple diagram of a Helium-4 atom

- Protons and neutrons are located in the nucleus at the center of the atom.
- Electrons orbit the nucleus in the electron cloud.
- Most of the mass of this helium atom is concentrated in the nucleus.
- The "4" is called the mass number - and refers to the total number of protons and neutrons in the nucleus.
- The protons, neutrons, and electrons is this illustration are much smaller than shown relative to the size of the atom. If an atom were the size shown here, you wouldn't be able to **see** the protons, neutrons, or electrons.

This picture of atoms is still simpler than the current view, but it will do for the moment. We will refine our picture of how the electrons are arranged around the atom near the end of the course.

The internal structure of the atom

The model of the atom in illustration 2 describes an atom as being made up of an **electron cloud** and a **nucleus**. The nucleus is made up of protons and neutrons. How do we describe the atoms?

The **chemical identity** of the atoms is set by the number of protons. The number of protons in the atomic nucleus is called the **atomic number**. Helium has an atomic number of 2; and each helium atom contains two protons. **All** helium atoms have two protons in their nucleus. A neutral (which you can read as electrically neutral) atom also has **a number of electrons equal to the atom's atomic number**. Thus, a neutral helium atom has two electrons.

The **mass** of the atom, represented by the **mass number** is determined by the number of protons **and** neutrons in the atom's nucleus. Electrons don't count here, as they weigh so little compared to protons and neutrons that they have a negligible effect on the mass. We call the species in illustration 2 "Helium-4". The "4" is the mass number - the total number of protons (2) plus the total number of neutrons (2).

Some atoms in nature have the same number of protons (thus the same element) but different numbers of neutrons. This gives them slightly different masses. These atoms of the same element with slightly different masses are called **isotopes**. Most elements have one or more isotopes. Hydrogen (the simplest element, with only one proton and one electron) has three isotopes. Normal hydrogen, Hydrogen-1 has one proton and no neutrons. Hydrogen-2, known as "deuterium", has one proton and one neutron. Hydrogen-3, known as "tritium", has one proton and two neutrons. Tritium is an unstable isotope of hydrogen. It's radioactive and used in the manufacture of nuclear weapons.

As a form of shorthand, nuclear chemists will often use the element's symbol and mass number to represent an element. For example, here is a table showing the representations of several **nuclides** (specific isotopes).

Nuclide	***Symbol***	***# Protons***	***# Neutrons***
Hydrogen-1	^{1}H	1	0
Hydrogen-2	^{2}H	1	1
Hydrogen-3	^{3}H	1	2
Helium-4	^{4}He	2	2
Carbon-12	^{12}C	6	6
Carbon-13	^{13}C	6	7

You may also see the atomic number as a subscript in front of the atomic symbol as well.

We should note here that **isotopes of a particular element behave identically in**

chemical reactions. For example, two atoms of hydrogen attach to one atom of oxygen to make water, H_2O. The same is true for deuterium and tritium. The only noticeable differences between hydrogen, deuterium, and tritium are their masses and the fact that tritium is radioactive. Chemically, they behave the same way.

<u>The weight of an atom</u>

We just mentioned that the isotopes of a different element differ by mass. We also talked about the mass number. Since a number is meaningless without a corresponding unit, what's the unit of the mass number?

The mass unit we use for things the size of atoms is the (appropriately named) **atomic mass unit**, or **amu**. This unit is defined based on the carbon-12 (^{12}C) isotope of carbon:

1 atom of ^{12}C has a mass of exactly **12.0000 amu**

Masses of other atoms are defined relative to carbon-12.

How do we find atomic masses? One way is through the use of an instrument called a **mass spectrometer**. The mass spectrometer is essentially an atom sorter that sorts atoms by mass. This can be done several ways, but the simplest (in concept) instrument available to do this is a **time of flight mass spectrometer**. Here's how it works:

1. **Positively-charged ions** are created in one end of the mass spectrometer. There are quite a few methods available for doing this. (An ion is simply an atom with an overall charge. Positively charged ions are atoms that have lost one or more electrons.)
2. These ions are accelerated by a negatively-charged grid and launched into a **flight tube**. The tube contains nothing - not even air. The **speed of the ions is determined by their respective weights**. The **smaller the ions are**, the **more speed they pick up** from the negatively charged grid. (Think of your ability to throw a baseball versus a bowling ball - you can throw a baseball a lot faster because it weighs less)
3. The ions hit a detector at the other end of the flight tube and are recorded. The larger the ion, the longer it takes for the ion to travel down the flight tube.

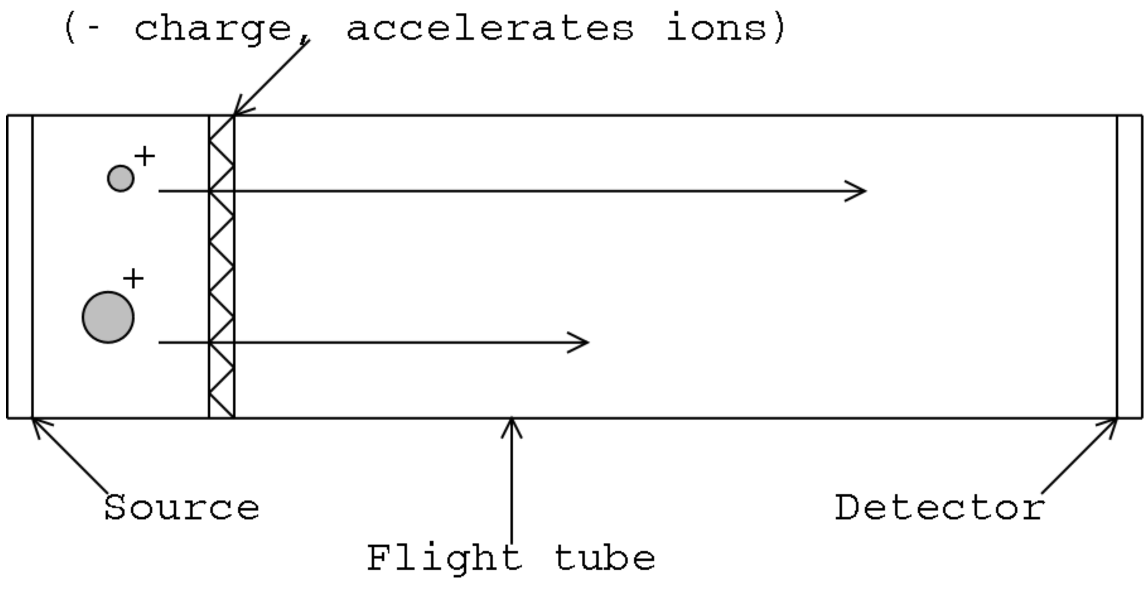


Illustration 3 - Simple time of flight mass spectrometer

With an instrument like a mass spectrometer, you can classify atoms by mass and record masses for different isotopes.

In nature, **each element exists as a mixture of isotopes**. Chlorine, for example, exists as a mixture of 76% ^{35}Cl and 24% ^{37}Cl (a form with two additional neutrons). The weight reported in the chemical literature is a weighted average of the masses of each isotope:

Isotope	***Mass Number****	***Percentage***	***Weighted Average Calculation***
^{35}C	35	76%	$35\times76\%=26.6\,\text{amu}$
^{37}C	37	24%	$37\times24\%=8.9\,\text{amu}$
Total (Both isotopes)	n/a	n/a	**35.5 amu**

If you look up the mass of naturally occurring chlorine, you'll find 35.5 amu as the reported mass. This is because 76% of the chlorine is ^{35}Cl, and 24% of it is ^{37}Cl.

**A more accurate number to use here would be the exact isotope mass, which is slightly different from the mass number*

<u>Summary</u>

In this note pack, we've discussed Dalton's atomic theory and the more modern nuclear model of the atom. Both of these theories give us a useful picture of how atoms are put together. We'll use this information later on to describe how atoms react with each other to form chemical compounds.