Elements – you will need to memorize the names and symbols of the following elements:

- Aluminum (Al)
- Uranium (U)
- Nitrogen (N)
- Nickel (Ni)
- Neon (Ne)
- Manganese (Mn)
- Lithium (Li)
- Hydrogen (H)
- Helium (He)
- Cobalt (Co)
- Chromium (Cr)
- Chlorine (Cl)
- Calcium (Ca)
- Cadmium (Cd)
- Bromine (Br)
- Boron (B)
- Bismuth (Bi)
- Barium (Ba)
- Arsenic (As)
- Antimony (Sb)

Common Elements

- Elements – you will need to memorize the names and symbols of the following elements:

Democritus

The Greek philosopher Democritus first proposed the idea of the atom back in 460 B.C. Democritus postulated that all matter could be subdivided until some finite particle was reached. This finite particle he called the atom, from the Greek word atomos, which means “indivisible”. Unfortunately, this idea was in conflict with the views of the philosopher Aristotle. Aristotle held the view that matter was a continuous substance which could not be divided into a fundamental unit. Since Aristotle was more widely known and held in greater esteem, it was the Aristotelian view of matter that predominated thinking into the first thousand years.

John Dalton

In the period 1803-1807 an English school teacher and chemist, John Dalton, developed an atomic theory of matter based on experimental evidence. Dalton postulated the existence of a different kind of atom for each element. He also postulated that atoms unite in definite ratios to form compounds and that these atoms joined in definite whole number ratios. The postulates of his theory can be summarized as follows:

- All matter is composed of atoms.
- All atoms of the same element are identical. Those of different elements are different.
- Atoms of one element cannot be converted into atoms of another element.
- Atoms unite in definite ratios to form compounds.

Dalton based his theory on three laws which had been formulated during the late 1700’s. By this time chemists had begun using quantitative methods in their experiments. These laws are summarized as follows:

1. The Law of Multiple Proportions

If two or more different compounds composed of the same two elements are analyzed, the mass percentage of elements as 79% tin and 21% oxygen:

- Tin (118.7 g), the ratio of oxygen in the tin oxide is 2:1.
- Tin (134.7 g), the ratio of oxygen in the tin oxide is 3:2.

Thus, the proportion by mass of the elements in a pure compound is a constant. For example:

\[ \text{Sn} + \text{O}_2 \to \text{SnO}_2 \]

The Law of Definite Proportions

The proportion by mass of the elements in a pure compound is a constant. For example:

\[ \text{Cu} + \text{S} \to \text{CuS} \]

The Law of Mass Conservation

Mass is neither created nor destroyed in any ordinary chemical reaction. Antoine Lavoisier first proposed the law when he observed the elements combined in inorganic compounds, mass was always conserved. Lavoisier then extended the law to organic compounds and determined the mass of the product was equal to the sum of the masses of the reactants.

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Cathode rays are capable of imparting mechanical motion. From this evidence it was concluded that cathode rays are particles of some kind.

J.J. Thomson discovered the particles comprising the cathode rays are negatively charged. He showed that if a magnet is placed outside a Crooke’s tube, the rays would always be deflected by the magnetic field in the direction that a negatively charged particle would be deflected. Thomson called the cathode ray particle an electron.

Thomson measured the charge/mass ratio of these electrons and found this value to be the same regardless of what gas was in the tube or what the electrodes were made of. Other experiments with Crooke’s tubes provided evidence for the existence of a fundamental unit of positive matter with a much greater mass, which was termed the proton. Thomson proposed a “Plum Pudding” model of the atom—the atom consisted of an equal mix of positive pudding and negative plums. This was the first model of the atom with subatomic particles.

Robert Millikan

In 1909 a professor at the University of Chicago, Robert Millikan, was able to determine the charge on an electron. He did this by way of the classic oil drop experiment. Based on Thomson’s charge/mass ratio of the electron, now the mass of the electron could be calculated. With this, the fundamental unit of negative charge in the atom, the electron, had now been characterized in terms of mass and charge. How did the oil drop experiment work?
Zapping the oil drops with X-rays caused them to become ionized – that is, the energy of the X-ray packed enough “punch” to knock off electrons of the oil drops causing them to become positively charged. The electrons were picked by other oil drops causing them to become negatively charged. The positively charged oil drops fell to the negative plate (opposites attract) while the negatively charged oil drops were repelled by the negative plate. By adjusting the electric force on the plates, Millikan could get an oil drop to be stationary (which he could observe through the “microscope”). At this point, the electric force (acting to push the drop up) = the gravitational force (acting to bring it down). That is,

\[ \text{Force}_{\text{electric}} = \text{Force}_{\text{gravity}} \]

Millikan knew all the variables in the equation except for the charge on the oil drop, which he solved for. This was the charge of the electron.

http://www.youtube.com/watch?v=XMfYHag7Liw

Ernst Rutherford

Ernst Rutherford performed the classic Gold Foil experiment with graduate students Ernest Marsden and Hans Geiger that led Rutherford to propose a nuclear model of the atom. This occurred between 1908 and 1911. In this experiment, Rutherford bombarded a thin sheet of gold foil with alpha particles. Alpha particles are a 2-proton and 2-neutron unit (a helium nucleus) that is emitted from the nuclei of unstable atoms. These alpha particles thus carry a net positive charge (2+). The set up is illustrated below:

Observations:
- Most alpha particles passed through
- A few alpha particles were deflected
Conclusions:
• Atoms are mostly empty space – most alpha particles passed through the gold atoms.
• Alpha particles deflected came close to concentrations of positive charge (like charges repel). The center of positive charge was termed the atom’s nucleus.

Rutherford observed similar results using other nuclei. The metals, however, needed to be extremely thin. Rutherford proposed that the atom consists of a tiny positively charged center that contains the mass of the atom which he termed the nucleus. This nucleus was surrounded by electrons which essentially accounted for the volume of the atom. This model of the atom is often called the planetary model of the atom. An electrostatic force (electrostatic forces are the attractive and repulsive forces between electrically charged particles). An electrostatic force of attraction keeps the electron “orbiting” around the nucleus while a “centrifugal” force keeps the electron from falling into the nucleus (and “collapsing” the atom). This presented a perfect explanation for the subatomic structure of the atom, but the model was short-lived.

http://www.youtube.com/watch?v=5pZj0u_XMbc

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**Isotopes of Hydrogen:**

- **Protium (H-1)**
- **Deuterium (H-2)**
- **Tritium (H-3)**

---

**Determining Atomic Masses**

The atomic masses listed for the elements on the periodic table are weighted averages of their isotopes. That is why their masses are shown with decimal places (masses of atoms). The number and percentage of each isotope of an element is determined by a mass spectrograph. This instrument is illustrated below:

---

<table>
<thead>
<tr>
<th>Particle</th>
<th>Relative mass</th>
<th>Relative charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>1 amu*</td>
<td>+1</td>
</tr>
<tr>
<td>Neutron</td>
<td>1 amu</td>
<td>0</td>
</tr>
<tr>
<td>Electron</td>
<td>1/1864 amu*</td>
<td>-1</td>
</tr>
</tbody>
</table>
A sample of an element is injected into the instrument, heated to vaporize it, and then bombarded with a beam of electrons. The electrons collide with atoms of the element knocking off some of the outermost electrons. The result is a series of atoms that have a net positive charge - positive gaseous ions. These ions are then accelerated by an electric field to a narrow stream and subjected to the force of a magnetic field. The ions are deflected by the magnetic field and separation occurs on the basis of mass, with the heaviest isotopes (those with the most neutrons) deflected the least while the lightest isotopes (with fewer neutrons) deflected the most. The atomic masses of the elements are based relative to the deflections of a carbon-12 isotope in a mass spectrometer. The C-12 isotope is defined as exactly 12 atomic mass units, or amu. Thus, 1 amu = 1/12 the mass of a C-12 atom.

The atomic mass of an element is computed from the masses of its isotopes and their fractional abundances. This is done by multiplying each isotope’s fractional abundance by its atomic mass, and then adding the results:

\[
\text{atomic mass} = (\text{atomic mass of isotope}_1 \times \%\text{ abundance}) + (\text{atomic mass of isotope}_2 \times \%\text{ abundance}) + (\text{atomic mass of isotope}_3 \times \%\text{ abundance}) + \text{etc.}
\]

**Problem:**

The element silicon consists of three isotopes: Si-28 (27.977 amu, 92.21%), Si-29 (28.976 amu, 4.70%), and Si-30 (29.974 amu, 3.09%).

a) Why can’t we simply add the masses of all three isotopes and divide by three to get the average mass?

b) To which isotope will the average mass be closest to (Si-28, Si-29, or Si-30)? Explain.

c) Calculate the atomic mass of silicon as reported on the Periodic Table.

d) Out of 100 atoms of a sample of silicon, how many will be the Si-28 isotope?

e) Out of 225 atoms of silicon, how many will be the Si-30 isotope?

f) Out of 100 atoms of silicon, how many atoms of silicon will have a mass of 28.086 amu?

**Solution:**

The atomic mass for chlorine is reported as 35.45 amu. It consists of two isotopes, Cl-35 (34.969 amu) and Cl-37 (36.966 amu). Calculate the percent (fractional abundance) of each isotope.

\[
\begin{align*}
\text{Cl-35} & \quad 92.12/100 = 0.9212 \\
\text{Cl-37} & \quad 8.74/100 = 0.0874
\end{align*}
\]

Out of 225 atoms of chlorine, how many will be the Cl-35 isotope?

\[
\frac{225 \times 0.9212}{100} = 219.62
\]

Out of 225 atoms of chlorine, how many will have a mass of 28.086 amu?

\[
\begin{align*}
\text{Cl-35} & \quad 219.62 \\
\text{Cl-37} & \quad 225 - 219.62 = 5.38
\end{align*}
\]

Use the following information to calculate the percent (fractional abundance) of each isotope.

\[
\begin{align*}
\text{The atomic mass for chlorine is reported as 35.45 amu. It consists of two isotopes, Cl-35 (34.969 amu) and Cl-37 (36.966 amu). } \\
\text{Calculate the percent (fractional abundance) of each isotope.}
\end{align*}
\]

Let \( x = 9.6 \text{Cl-35} \)

Then \( 1 - x = 9.4 \text{Cl-37} \)

\[
\begin{align*}
35.453 &= (35.45 + x) \\
&= (35.45 + 9.6) \\
&= 35.45 + 35.45 + 9.6 \text{Cl-35} \\
&= 35.45 + 29.11 \\
&= 35.45 + 29.11 + 9.6 \text{Cl-35}
\end{align*}
\]

\[
\begin{align*}
X &= 75.979 \text{Cl-35} \\
&= 100 - 75.979 \\
&= 24.021
\end{align*}
\]
The Periodic Table

The elements known to the ancient world included iron (Fe), copper (Cu), silver (Ag), gold (Au), mercury (Hg), tin (Sn), lead (Pb), carbon (C) and sulfur (S). The alchemists “discovered” other elements that included cobalt (Co), nickel (Ni), zinc (Zn), antimony (Sb), arsenic (As), and phosphorous (P). These were the elements known up to the 18th century. Credit for the discovery of the modern Periodic Table is generally given to the chemist Dmitri Mendeleev (1869). Mendeleev arranged the known elements on the basis of the atomic masses in rows (periods) in such a way that elements with similar properties fell into the same vertical columns (groups or families). He recognized the combining capacities of elements as a fundamental classifying characteristic and considered the periodic arrangement as a natural law with predictive powers. He placed iodine (I) after tellurium (Te), for example, because it had similar properties with chlorine (Cl) and bromine (Br) despite the fact that it had a smaller atomic mass. The key to Mendeleev’s Periodic Table, and the reason he is acknowledged as the discoverer of the Periodic Table, was that he recognized that elements may not yet have been discovered. He left blanks in his Table where he seemed to be missing, and he very accurately predicted the properties of the yet undiscovered elements gallium (Ga), scandium (Sc) and germanium (Ge). Consider the case of germanium as an example:

<table>
<thead>
<tr>
<th>Atomic mass</th>
<th>Density</th>
<th>Color</th>
<th>Density of oxide</th>
<th>Formula of chloride</th>
<th>Formula of oxide</th>
</tr>
</thead>
<tbody>
<tr>
<td>GeCl4</td>
<td>72.6</td>
<td>grey-white</td>
<td>5.47</td>
<td>GeO2</td>
<td>GeO2</td>
</tr>
<tr>
<td>GeO2</td>
<td>72.6</td>
<td>grey-white</td>
<td>5.47</td>
<td>GeCl4</td>
<td>GeO2</td>
</tr>
<tr>
<td>Ge</td>
<td>72.6</td>
<td>grey-white</td>
<td>5.47</td>
<td>GeO2</td>
<td>GeO2</td>
</tr>
</tbody>
</table>

Modern Periodic Table: the physical and chemical properties of the elements are periodic functions of their atomic numbers. Arrangement of Table:

- Periods: the number of each period indicates the principle energy level in which the outermost or valence electrons of that period’s elements are found. In each period the number of valence electrons increases from left to right.
- Groups: elements within a group exhibit similar or related properties because they have the same number of valence electrons.

<table>
<thead>
<tr>
<th>Group</th>
<th>Elements</th>
<th>Properties</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>alkali metals</td>
<td>poor conductors of heat and electricity, malleable, ductible, silvery-white, lustrous, react vigorously with water (H2) and nonmetallic oxides (e.g., CO2)</td>
</tr>
<tr>
<td>2</td>
<td>alkaline earth metals</td>
<td>reactive with water, resistant to many nonmetallic oxides</td>
</tr>
<tr>
<td>12</td>
<td>transition metals</td>
<td>reactive with water, resistant to many nonmetallic oxides</td>
</tr>
<tr>
<td>18</td>
<td>noble gases</td>
<td>inert, do not react with water or nonmetallic oxides</td>
</tr>
</tbody>
</table>

Metals are excellent conductors of heat and electricity; metallic bonding is such that the atoms of metallic atoms (lattice = positive ions) are held in fixed positions in a crystalline arrangement and the valence electrons are shared by the entire crystal. Metals are malleable, ductile, good conductors of heat and electricity; metallic bonding is such that the kernels of metallic atoms (pycnoclasts + inner electrons) are held in fixed positions in a crystalline arrangement and the valence electrons are shared by the entire crystal.
Metalloids (semi-metals)
- elements that exhibit both metallic and nonmetallic properties; these include B, C, Ge, Si, As, and Te.

Noble Gases
- have eight valence electrons (except He which has 2).
- chemically very unreactive although heavier krypton (Kr) and xenon (Xe) have formed compounds by reacting with the highly electronegative elements fluorine and oxygen under extreme conditions.

Allotropy
- elements can occur in different forms. These different forms are termed allotropes.
- There are three allotropes of the element carbon:
  - diamond – carbon atoms occur in a tetrahedral arrangement
  - graphite – carbon atoms are bonded in a layered structure with weak forces of attraction between the layers
  - buckminsterfullerenes – 60 carbon atoms bonded in a polyhedral structure

Oxygen occurs in two allotrope forms:
- oxygen as a diatomic molecule, O₂ (called oxygen)
- ozone, which is a triatomic molecule – O₃. Ozone occurs in our upper atmosphere and is critically important in absorbing the ultraviolet rays released by our sun. Ozone can also be found near the ground after it is made during lightning storms. Near the ground ozone is a health risk.

Ions
- The Rutherford model of the atom we have discussed consists of a nucleus containing the protons and neutrons surrounded by electrons. In a neutral atom, the number of protons (positives) equals the number of electrons (negatives). Thus, a neutral atom is said to have zero net charge.
- We will learn later that when elements react with each other to form compounds they do so with their outermost or valence electrons. The number of electrons that occur in each energy level of an atom is given by its "electron configuration", which appears on your periodic table in the lower left corner of the element's box. Sodium, for example, has the electron configuration 2, 8, 1, which means it has 2 electrons in level 1, 8 electrons in level 2, and 1 electron in level 3.
- Recall we said that it is the valence electrons of an element that determine its properties and how it reacts with other elements. The model that chemists have developed to use our theory of chemical bonding is that elements will gain, lose or share electrons in bonding with other elements. For now, we will be interested in how elements lose or gain electrons. In general, metals tend to lose electrons when reacting with nonmetals to form positively charged ions called cations. Nonmetals tend to gain electrons when reacting with metals to form negatively charged ions called anions. Thus, ions (atoms with a net electric charge) are formed when atoms lose or gain electrons.

Let's consider the formation of simple cations and anions:

Potassium → potassium ion
\[ \text{K}^+ \]

Fluorine → fluoride ion
\[ \text{F}^- \]

19 protons (+) \[ \rightarrow \]
19 electrons (−)

net charge = 0

k+ \[ \rightarrow \]

10 electrons (−)

net charge = −1

A cation is named using the name of the parent. Thus, K⁺ is called potassium ion. An anion is named by taking the root of the name and changing the ending to "ide". Thus, F⁻ is called fluoride ion.

The figure below shows the types of ions formed by atoms in several groups on the periodic table:
Ionic Compounds

The chemist Svante Arrhenius was the first to postulate that chemical compounds could be made up of oppositely charged particles. In his research, Arrhenius observed that neither pure water nor a dry salt would conduct an electric current. However, if the salt was dissolved in water, the solution would conduct an electric current. He explained his observations by suggesting that units of salt broke up into charged particles or "ions" when dissolved in water and that explained the conductivity of the solution. The idea of "ions" was pretty radical back in the 1800s (no pun intended) and Arrhenius' idea was not accepted. His peers argued that soluble chloride dissolved in water gives no evidence of a greenish chlorine gas which they said should accompany the decomposition of sodium chloride. Arrhenius was correct to point out that chloride ion has different properties compared to the chlorine atom. His idea eventually became widely accepted and in 1903, Arrhenius was awarded the Nobel Prize in chemistry.

Today we know that in order for a substance to conduct an electric current one of two criteria must be satisfied:

- The substance must contain freely moving electrons (as occurs in metals)
- The substance must have freely moving ions. Compounds consisting of ions meet this requirement in the liquid (or molten) and aqueous (dissolved in water) phases.

Ionic compounds meet the second criteria. Ionic compounds form when a metal reacts with a nonmetal. In the reaction, the metal transfers an electron(s) to the nonmetal resulting in a positively charged metal ion and a negatively charged nonmetal ion. It is this electrostatic attraction between oppositely charged ions that results in the ionic bond. More about bonding later!

In writing the formulas of ionic compounds, the net charge must be conserved. Cations and anions will react in a ratio that conserves charge. When writing the formulas of ionic compounds, we adjust the subscripts to obtain the correct ratio. In addition, we typically express the chemical formula of an ionic compound by using the simplest whole number ratio of atoms. This is referred to as the empirical formula. We have discussed the common charges for nonmetals. For metals, the Group 1 alkali metals all have charges of +1 only while that of the Group 2 alkaline earth metals all have charges of +2 only.

Let's try the following examples involving simple binary ionic compounds using the "criss-cross" rules:

- Sodium bromide: $K^+ + Br^- \rightarrow KBr$
- Lithium oxide: $Li^+ + O^{2-} \rightarrow Li_2O$
- Magnesium fluoride: $Mg^{2+} + F^- \rightarrow MgF_2$
- Calcium oxide: $Ca^{2+} + O^{2-} \rightarrow CaO$
- Calcium nitride: $Ca^{2+} + N^{3-} \rightarrow Ca_3N$
- Aluminum oxide: $Al^{3+} + O^{2-} \rightarrow Al_2O$
- Barium sulfide: $Ba^{2+} + S^{2-} \rightarrow BaS$
- Rubidium sulfide: $Rb^{+} + S^{2-} \rightarrow Rb_2S$
- Aluminum triiodide: $Al^{3+} + I^{-} \rightarrow AlI_3$

- $Ba^{2+} + S^{2-} \rightarrow BaS$
- $Rb^{+} + S^{2-} \rightarrow Rb_2S$
- $Al^{3+} + I^{-} \rightarrow AlI_3$