Atoms, Molecules, Formula, and Subatomic Particles

Atoms

Atoms are the immensely small particles of which all matter is composed. How small are atoms? They are so small that they have only been able to be seen in the last couple of decades. A regular microscope will not resolve atoms or molecules, the next step up in the ladder of size. Atoms require either a scanning tunneling microscope or an atomic force microscope to see. These microscopes do not use light. Instead, they use the forces between the atoms to “see” the atoms or molecules themselves. Atoms are particles that, now that we can see them, we can measure them, are approximately 10^{-10} m is size. The mass of an atom is approximately 10^{-23} g.

The idea of atoms is not a new one. It dates back to the Greek philosophers Democritus and Leucippus. They hypothesized that all matter must be made up of very tiny indestructible particles called atoms. The word atom comes from the Greek for indivisible. A couple of millennia later the English chemist John Dalton revived this old idea and added a little more to it. He developed the **atomic theory of matter** by making 5 statements that describe the theory. The statements are:

1. All matter is composed of small particles called atoms. A type of atom corresponds to a different element.
2. All atoms of a given element are similar to one another and very different from other elements.
3. The relative number and arrangement of the elements contained in a pure substance determines its identity.
4. A chemical change is a rearrangement of the atoms to give new substances.
5. Only whole atoms can participate in a chemical change.

Molecules

Molecules are collections of atoms that are chemically bound together. The group of atoms, because they are chemically bound together, act as a single unit. We can image molecules using the same technology we use to image atoms. The molecules themselves are extremely tiny, albeit larger than most atoms. The smallest kind of molecule is composed of only two other atoms. The atoms can be of the same kind or of different kinds. Molecules that are composed of two atoms of the same kind are also referred to as elements. All molecules that have only two atoms are called diatomic molecules. There are molecules of varying complexity from the diatomic molecules up to molecules like DNA, which contains approximately 20 billion atoms.

Molecules can be composed of only one kind of atoms, which we stated earlier were called elements. These molecules are also called **homoatomic** molecules. If the molecule contains more than one kind of atom, it is called **heteroatomic**. The presence of homoatomic molecules indicates that not all elements prefer to be in the state of individual atoms. There are elements that are molecular and composed of two, four, or even eight atoms. The most common diatomic molecular elements are Hydrogen, Nitrogen, Oxygen, Fluorine, Chlorine, Bromine, and Iodine. Phosphorus exists as a molecule with four atoms. Sulfur is a ring of eight atoms. ***You should memorize the elements that are diatomic molecules***.

Not all compounds are molecules. Some compounds are collections of positively and negatively charged atoms called ions. These are called **ionic compounds**. For molecular compounds, the smallest unit is the molecules. For ionic compounds, the smallest unit is the formula unit, which is the lowest whole number ration of the ions present.

The molecule is the smallest part of a molecular compound that retains the chemical and physical properties of the compound. For an ionic compound, it is the formula unit. Some examples of elements and compounds are listed below:

Formula	Type
CH_4	molecular compound
NaCl	ionic compound
H_2O	molecular compound
P_4	tetratomic molecular element
N_2	diatomic molecular element

Natural vs. Synthetic compounds

There is some discussion as to whether natural compounds are better for you than synthetic compounds. For example, is the vitamin C in orange juice better than the vitamin C made in a lab? The answer to this is no. There is no difference between the molecules of vitamin C in the OJ you drink at breakfast and the molecules of vitamin C that you take in your vitamin supplement, most likely produced in a lab. If the molecule has the same arrangement of the same kinds of atoms, it is the same kind of molecule regardless of where or how it is made. There are claims by companies that market “natural” supplements that theirs are better than the “synthetic” supplements. This is merely a marketing ploy. They believe, and in some cases, rightly so, that people will be willing to pay more for the natural variety. People need to realize that the compounds in both are the same. Only the perceived difference is the reason for the increased cost.

Chemical Formulas (Formulæ)

Chemical formulas are a shorthand method of representing chemical compounds. The formula represents the number of each type of atoms present in the compound. For example, H_2O means that the molecule, water, contains 2 atoms of hydrogen and 1 atom of oxygen.

When writing chemical formulas you need to be aware of the capitalization rules for the symbols of the elements. $CoCl_2$ and $COCl_2$ are very different compounds. NACL does not exist because there is no element that has either A or L as its symbol.

For molecular compounds, the formula is a direct indication of the numbers of atoms in the compound. For ionic compounds, the formula is the lowest whole number ratio of the ions present in the compound. The chemical formula of an ionic compound may contain parenthesis (e.g., $(NH_4)_2SO_4$). The parenthesis means that there are whatever number is outside of the parenthesis of what is inside. In the example above, there is 1 atom of nitrogen (1×2), 8 atoms of hydrogen (4×2), 1 atoms of sulfur and 4 atoms of oxygen.

Subatomic particles: protons, neutrons, and electrons, Oh! My!

All atoms are composed of at least 1 proton and one electron. Most atoms have at least one neutron. Evidence from the latter part of the 19th and the beginning of the 20th centuries indicated that the atom was not as indestructible and fundamental as was originally thought. The positively charged protons were discovered in 1886, the negatively charged electrons in 1897, and the neutral neutron in 1932. All of these particles were found to be within most atoms.

Protons are subatomic particles that have a positive charge and a mass of 1 atomic mass unit (amu) (1.673×10^{-24} g). **Electrons** have a negative charge equal to, but opposite, the charge of the proton and a mass of 0 amu (actually equal to 9.109×10^{-28} g). **Neutrons** have no charge and a mass that is approximately equal to that of the proton (1 amu, 1.675×10^{-24} g).

Arrangement of particles in the atom

All of the particles have their place in the atom. The protons and neutrons are contained in the nucleus of the atom. Almost all of the mass (99.999%) of an atom is contained in the nucleus. Because the nucleus contains only protons and neutrons, it is positively charged.

The electrons are arranged around the nucleus. Most of the space where the electrons are is empty. The motion of the electrons in this region determines the size of the atom.

In a neutral atom, the number of protons in the nucleus is equal to the number of electrons in the outer part of the atom. This means that all atoms have no charge.

The nucleus of the atom is very tiny compared to the atom itself. We saw earlier that the atom was approximately 10^{-10} m in diameter. The nucleus is approximately 10^{-15} m in diameter.

To get an idea of how concentrated the mass of the atom is in the nucleus, we can think about neutron stars. These stars have collapsed under their own mass and forced the electron and protons together to make neutrons. The neutron star is essentially a giant (approximately 10 km) atomic nucleus composed of only neutrons. One teaspoon of neutron star material has a mass equivalent to that of a mountain.

Evidence for the structure of the atom

You are not responsible for knowing these experiments in detail or who carried them out. You should know what each of the experiments tells us about the structure of the atom.

Discharge tube experiments

Discharge tubes are tubes that contain a very thin gas. These tubes evolved into today's neon lights and television monitors. The principle behind this was discovered in 1821 by Davy. In the late 1800's J.J. Thomson and Eugene Goldstein independently discovered electrons and protons, respectively. As a result of these experiments, the atom was believed to be somewhat like a plum pudding. The plums were the electrons in a pudding of positively charged protons.

Metal Foil experiments

In 1911 Ernest Rutherford designed an experiment that changed the way everyone thought about the atom. His experiment was set up to explore the validity of the above model. He bombarded gold foil with alpha particles. Alpha particles were known at the time (discovered in 1896) to be positively charged, but it was not known that they were helium nuclei. If the "plum pudding" model was correct, the alpha particles should mostly go through the foil with some being deflected slightly to the sides.

What Rutherford saw instead was, first, what he expected, and second, some alpha particles being bounced back toward the source. This was entirely unexpected. Rutherford said that is was like firing a 15 in gun at a piece of tissue paper and having the shell bounce back at you.

From this experiment, Rutherford was able to deduce the currently accepted structure of the atom. A problem scientists had with this model was that there was nothing to hold the positively charged protons together in the nucleus. This was solved in 1932 with the discovery of the neutron.

Atomic number and mass number

Atomic number is the number of protons in a nucleus. All atoms with the same atomic number are the same element. If an atom contains 67 protons in the nucleus, regardless of the number of electrons or neutron, it is a holmium atom. If there are 33 protons, it is an arsenic atom. We can now define an element as a pure substance in which all atoms present have the same atomic number. If it is a neutral atom, the number of electrons is also equal to the atomic number.

The **mass number** is the total number of protons and neutrons in the nucleus. This number gives an indication of the mass of the atom. It is NOT, however, the mass of the atom. The mass number is an integer, always. The mass of the atom is not an integer.

We can explicitly give the atomic number, Z, and the mass number, A, with the symbol. This is shown here:

$$ {}_{Z}^{A}E \,. $$

The atom of oxygen with 8 protons (by definition) and 9 neutrons is represented as:

$$ {}_{8}^{17}O \,. $$

Isotopes

All atoms of a given element must have the same number of protons. There is indicating how many neutrons an element must have. It must have enough neutrons to hold the protons together, but this number may vary. Hydrogen, for instance, can have zero, one or two neutrons in its nucleus. These different kinds of hydrogen atoms are called **isotopes**. Isotopes are atoms of the same element with different numbers of neutrons. The isotopes of hydrogen are represented as:

$$ {}_{1}^{1}H \qquad {}_{1}^{2}H \qquad {}_{1}^{3}H \,. $$

Isobars are atoms with the same mass number but different atomic numbers. As a result, isobars are different elements. Isotopes are more commonly used than isobars.

Atomic Masses

The masses in the periodic table are not mass numbers, in general. They are relative average atomic masses. Why are they relative? Because the masses are defined relative to something. That something is the isotope of carbon with a mass number of 12. This isotope is defined to have a mass of exactly 12.0000 amu. All other atomic masses are defined relative to this. The masses in the periodic table are also average masses. They are average masses because natural elements are composed of a mixture of different isotopes. For instance, Hydrogen is approximately 99% hydrogen with a mass number of 1 and 1% with a mass number of 2. The average mass is a weighted average. We weight the more prevalent isotope more heavily. In order to calculate the relative average atomic mass of an element we need to know

1. the relative masses of the individual isotopes, and
2. the abundances of the individual isotopes, and
3. the number of isotopes.

Example of calculation of relative average atomic mass of an element:

Silicon is composed of three naturally occurring isotopes (data taken from the 1999 CRC Handbook of Chemistry and Physics): Silicon-28 (${}^{28}_{14}Si$), silicon-29 (${}^{29}_{14}Si$), and silicon-30 (${}^{30}_{14}Si$). Silicon-28 occurs 92.22% of the time, silicon-29: 4.69% and silicon-30: 3.09%. The mass of the three isotopes is: silicon-28, 27.97692653 amu; silicon-29, 28.97649472 amu; and silicon-30, 29.97377022 amu. The following table summarizes this data:

Isotope	mass (amu)	% abundance
silicon-28	27.97692653	92.22
silicon-29	28.97649472	4.69
silicon-30	29.97377022	3.09

We calculate the relative average atomic mass by multiply the abundance ***as a fraction*** by the corresponding mass and adding these products together (the underlined digit is the last significant figure in that number).

$$(0.9222)(27.97692653\text{ amu})+(0.0469)(28.97649472\text{ amu})+(0.0309)(29.97377022\text{ amu})$$
$$=25.8\underline{0}032165\text{ amu}+1.3\underline{5}8997602\text{ amu}+0.092\underline{6}189499\text{ amu}=28.0\underline{8}550875\text{ amu}=28.09\text{ amu}$$