Properties According to the Periodic Table

Atomic Radius

Figure shows the relative sizes of the atoms of the representative elements. Notice that atom size increases from top to bottom in a column and from right to left across a row. This trend is related to electron configuration. As we look at the elements in column 1, for example, we see that the single valence electron for each successive element is in a higher principal energy level than the last, and the electron is thus farther away from the positively charged nucleus; hence, the atomic radius increases going from top to bottom. This same regular increase in size can be observed in each column of the periodic table.

Atoms decrease in size going across a period from left to right. For elements within a period, electrons are being added one by one to the same principal energy level. At the same time, protons are also being added one by one to the nucleus, increasing its positive charge. This increasing positive charge increases the attraction of the nucleus for all electrons and pulls them all closer to the nucleus, decreasing the atom's radius. Thus, atomic size is a periodic property that increases from top to bottom within a column and from right to left across a period.

Ionization Energy

The ionization energy of an element is the minimum energy required to remove an electron from a gaseous atom of that element, leaving a positive ion. An equation expressing the ionization of sodium would be:

$$Na \longrightarrow Na^{+} + e^{-}$$

Electrons are held in the atom by the attractive force of the positively charged nucleus. The farther the outermost electrons are from the nucleus, the less tightly they are held. Thus, the ionization energy within a group of elements decreases as the elements increase in atomic number. Among the atoms of naturally occurring alkali metals, the single valence electron of cesium is farthest from the nucleus (in the sixth principal energy level), and we can correctly predict that the ionization energy of cesium is the lowest of all the alkali metals. (Recall that francium is not naturally occurring.)

From left to right across a period, the ionization energy of the elements tends to increase. The number of protons in the nucleus (the nuclear charge) increases, yet the valence electrons of the elements are in the same energy level. It becomes increasingly more difficult to remove an electron from the atom. The ionization energy of chlorine is much greater than that of sodium, an element in the same period.

The ionization energies of elements 1 through 36 are plotted versus their atomic numbers in Figure 5.15. The peaks of the graph are the high ionization energies of the noble gases. The height of the peaks decreases as the number of the highest occupied energy level increases. The low points of the graph are the ionization energies of the alkali metals, which have only one electron in their valence shell. These points, too, decrease slightly as the number of the highest occupied energy level increases. The graph shows that ionization energy is periodically related to atomic number. Even within a row of the periodic table, the variations in ionization energy are closely related to electron configuration.

The ionization energy of an element is a measure of its metallic nature. From Figure 5.15, we see that each alkali metal has the lowest ionization energy of the elements in its period. Therefore, alkali metals are the most metallic elements. From bottom to top in the periodic table and from left to right across it, the metallic nature of the elements decreases.

Nonmetals, located in the upper-right section of the periodic table, have high ionization energies. Except for the noble gases, fluorine has the highest ionization energy. Therefore, excluding the noble gases, fluorine is the least metallic (or most nonmetallic) element. From top to bottom in a column or to the left of fluorine, elements become more metallic. In summary, ionization energy increases from bottom to top of a column and from left to right across a period.

Electron affinity is closely related to but the opposite of ionization energy. Electron affinity is the energy change that occurs when an electron is added to a neutral atom. For a nonmetal this change is usually a release of energy. The equation showing this reaction for chlorine is:

$$Cl_2 + 2\,e^- \longrightarrow 2\,Cl^-$$

Electron affinities are fairly difficult to measure. Accurate values have been determined for only a few elements. In general the values become increasingly negative from left to right across a period. Consequently, a halogen will have the most-negative electron affinity of all the representative elements in its period (remember that the noble gases are not representative elements). Electron affinity does not change with the same regularity as does atomic radius or ionization energy, and relative values cannot be predicted as easily.

Trends of various atomic properties as related to position in the periodic table.

The Formation of Ions

Atoms are electrically neutral. The number of positively charged protons in the nucleus of an atom equals the number of negatively charged electrons outside the nucleus. If electrons are added or lost as an atom reacts, the atom acquires a charge and becomes an ion.

The octet rule

We have already observed that the noble gases are very unreactive. This lack of reactivity is attributable to a stable electron configuration. Looking back to Section 5.5C, you can see that all the noble gases but helium have eight electrons (two s and six p) in the highest occupied energy level. When atoms of the other representative elements react, they lose, gain, or share enough electrons to attain the noble-gas electron structure - a complete octet, eight electrons, in their outer shell. This tendency is expressed by the octet rule: An atom generally reacts in ways that give it an octet of electrons in its outer shell. Hydrogen and lithium are exceptions; they react in ways that give them the same electron configuration as helium, with two outer-shell electrons.

An atom with one, two, or three valence electrons usually reacts by losing these electrons to acquire the electron configuration of the noble gas next below it in atomic number. An atom with six or seven valence electrons will usually react by adding enough electrons to acquire the electron configuration of the noble gas next above it in atomic number. Other atoms may attain complete octet by sharing electrons with a neighboring atom

Positive ions, or cations

When a neutral atom loses an electron, it forms a positively charged ion, called a cation (pronounced "cát-i-on"). In general, metals lose electrons to form cations. The atom thereby attains the electron configuration of the noble gas next below it in atomic number.

For example, an alkali metal loses one electron to form a cation with a single positive charge. Sodium loses its single 3s valence electron to form the ion Na+, which has the electron configuration of neon:

$$\mathrm{Na}\cdot \longrightarrow \mathrm{Na}^{+} + \mathrm{e}^{-}$$

An alkaline earth metal loses two electrons to form a cation with a charge of +2. In forming the magnesium ion, Mg2+, a magnesium atom loses its two valence electrons:

$$\mathrm{Mg}\colon \longrightarrow \mathrm{Mg}^{2+} + 2\,\mathrm{e}^{-}$$

Aluminum loses its three valence electrons to form a cation with a charge of +3:

$$\mathrm{Al}\colon \longrightarrow \mathrm{Al}^{3+} + 3\,\mathrm{e}^{-}$$

Transition elements and the metals to their right do not always follow the octet rule; frequently they form more than one cation. For example, iron forms Fe2+ and Fe3+; cobalt forms Co2+ and Co3+. The names of these ions must indicate the charge they carry. The preferred system of nomenclature (naming) is that recommended by the International Union of Pure and Applied Chemistry (IUPAC). In this system, the name of the metal is followed by a Roman numeral (in parentheses) showing the charge on the ion. No extra space is left between the name and the number. Thus, Fe2+ is iron(II) (pronounced "iron two"), and Fe3+ is iron(III). In the older system, the name of the cation of lower charge ends in ous, and the name of the cation of higher charge ends in ic.

Negative ions, or anions

When a neutral atom gains an electron, it forms a negatively charged ion, called an anion (pronounced "án-i-on"). Typically, nonmetals form anions, gaining enough electrons to acquire the electron configuration of the noble gas of next higher atomic number. Elements of group 6, with six valence electrons, form anions by gaining two electrons; the halogens, with seven valence electrons, form anions by gaining one electron. The names of these anions include the root name of the element and the ending ide.

Polyatomic ions

The ions described in the preceding paragraphs are monatomic ions; that is, each contains only one atom. Many polyatomic ions are also known. Polyatomic ions are groups of atoms bonded together that carry a charge due to an excess or deficiency of electrons. The symbols in the formula show which elements are present. The subscripts ("1" is understood) tell how many atoms of each element are present in the ion.

Mole & Atomic Structure

An atom is very small. Its mass is between 10-21 and 10-23g. A row of 107 atoms (10,000,000 atoms) extends only 1.0 mm. We know that atoms contain many different subatomic particles such as electrons, protons, and neutrons, as well as mesons, neutrinos, and quarks. The atomic model used by chemists requires knowledge of only electrons, protons, and neutrons, so our discussion is limited to them.

The Electron

An electron is a tiny particle with a mass of 9.108 X 10-28g and a negative charge. All neutral atoms contain electrons. The electron was discovered and its properties defined during the last quarter of the nineteenth century. The experiments that proved its existence were studies of the properties of matter in gas-discharge or cathode-ray tubes.

Each electron carries a single, negative electric charge and has a mass of 9.108 X 10-28g. Because the mass of an atom is approximately 10-23g, the mass of an electron is negligible compared to that of an atom.

The Proton

Gas-discharge tubes of slightly different design were used to identify small, positively charged particles that moved from the positive electrode (anode) to the negative electrode (cathode). The mass and charge of these particles varied but were always a simple multiple of the mass and charge of the positive particle observed when the gas-discharge tube contained hydrogen. The particle formed from hydrogen is called the proton. The mass of a proton is 1.6726 X 10-24g, or about 1836 times the mass of an electron. The proton carries a positive electrical charge that is equal in magnitude to the charge of the electron but opposite in sign. All atoms contain one or more protons.

The Neutron

The third subatomic particle of interest to us is the neutron. Its mass of 1.675 X 10-24g is very close to that of the proton. A neutron carries no charge. With the exception of the lightest atoms of hydrogen, all atoms contain one or more neutrons.

Particle	Actual mass (g)	Relative mass (amu)	Relative charge
proton	1.6726×10^{-24}	1.007	+1
neutron	1.6749×10^{-24}	1.008	0
electron	9.108×10^{-28}	5.45×10^{-4}	-1

Because the actual masses of atoms and subatomic particles are so very small, we often describe their masses by comparison rather than in SI units, hence the term relative mass. If a proton is assigned a mass of 1.007, then a neutron will have a relative mass of 1.008 and an electron a mass of 5.45 X 10-4. When talking about relative masses, we use the term atomic mass unit (amu). Using this unit, a proton has a mass of 1.007 amu, a neutron a mass of 1.008 amu, and an electron a mass of 5.45 X 10-4 amu. Charges, too, are given relative to one another. If a proton has a charge of +1, then an electron has a charge of -1

Atomic Structure

Atomic Number Equals Electrons or Protons

Each element has an atomic number. The atomic numbers are listed along with the names and symbols of the elements on the inside cover of the text. The atomic number equals the charge on the nucleus. It therefore also equals the number of protons in the nucleus and also equals numerically the number of electrons in the neutral atom. The atomic number has the symbol Z. Different elements have different atomic numbers; therefore, atoms of different elements contain different numbers of protons (and electrons). Oxygen has the atomic number 8; its atoms contain 8 protons and 8 electrons. Uranium has the atomic number 92; its atoms contain 92 protons and 92 electrons. The relationship between atomic number and the number of protons or electrons can be stated as follows:

Atomic number	**= number of protons per atom**
	= number of electrons per neutral atom

Mass Number Equals Protons plus Neutrons

Each atom also has a mass number, denoted by the symbol A. The mass number of an atom is equal to the number of protons plus the number of neutrons that it contains. In other words, the number of neutrons in any atom is its mass number minus its atomic number.

Number of neutrons = mass number - atomic number

or

Mass number = number of protons + number of neutrons

The atomic number and the mass number of an atom of an element can be shown by writing, in front of the symbol of the element, the mass number as a superscript and the atomic number as a subscript:

	mass number atomic number	Symbol of element	or	A Z	**X**

For example, an atom of gold (symbol Au), with an atomic number 79 and mass number of 196 is denoted as:

196 79	Au

Isotopes

Although all atoms of a given element must have the same atomic number, they need not all have the same mass number. For example, some atoms of carbon (atomic number 6) have a mass number of 12, others have a mass number of 13, and still others have a mass number of 14. These different kinds of the same element are called isotopes. Isotopes are atoms that have the same atomic number (and are therefore of the same element) but different mass numbers. The various isotopes of an element can be designated by using superscripts and subscripts to show the mass number and the atomic number. They can also be identified by the name of the element with the mass number of the particular isotope. For example, as an alternative to

	12 6	C	13 6	C	and	14 6	C

we can write carbon-12, carbon-13, and carbon-14. About 350 isotopes occur naturally on Earth, and another 1500 have been produced artificially. The isotopes of a given element are by no means equally abundant. For example, 98.89% of all carbon occurring in nature is carbon-12, 1.11% is carbon-13, and only a trace is carbon-14. Some elements have only one naturally occurring isotope.

The Inner Structure of the Atom

So far, we have discussed electrons, protons, and neutrons and ways to determine how many of each a particular atom contains. The question remains: Are these particles randomly distributed inside the atom like blueberries in a muffin, or does an atom have some organized inner structure? At the beginning of the twentieth century, scientists were trying to answer this question. Various theories had been proposed, but none had been verified by experiment. In our discussion of the history of science, we suggested that, at various points in its development, science has marked time until someone performed a key experiment that provided new insights. In the history of the study of atoms, a key experiment was performed in 1911 by Ernest Rutherford and his colleagues.

Forces between bodies

Our understanding of the conclusions drawn from Rutherford's experiment depends on a knowledge of the forces acting between bodies. Therefore, before discussing his experiment, a brief review of these forces is in order. First is the force of gravity that exists between all bodies. Its magnitude depends on the respective masses and on the distance between the centers of gravity of the two interacting bodies. You are familiar with gravity; it acts to keep your feet on the ground and the moon in orbit. Electrical forces also exist between charged particles. The magnitude of the electrical force between two charged bodies depends on the charge on each body and on the distance between their centers. If the charges are of the same sign (either positive or negative), the bodies repel each other; if the charges are of opposite sign, the bodies attract each other. Magnetic forces, a third type, are similar to electrical forces. Each magnet has two poles - a north pole and a south pole. When two magnets are brought together, a repulsive force exists between the like poles and an attractive force between the unlike poles. The magnetic and electrical forces can interact in the charged body. These three forces were known at the end of the nineteenth century when the structure of the atom came under intensive study.

Rutherford's experiment

Let us describe Rutherford's experiment, In 1911, it was generally accepted that the atom contained electrons and protons but that they were probably not arranged in any set pattern. Rutherford wished to establish whether a pattern existed. He hoped to gain this information by studying how the protons in the atom deflected the path of another charged particle shot through the atom. For his second particle, he chose alpha ( ) particles. An alpha particle contains two protons and two neutrons, giving it a relative mass of 4 amu and a charge of +2. An alpha particle is sufficiently close in mass and charge to a proton that its path would be changed if it passed close to the proton. In the experiment, a beam of alpha particles was directed at a piece of gold foil, so thin as to be translucent and, more importantly for Rutherford, only a few atoms thick. The foil was surrounded by a zinc sulfide screen that flashed each time it was struck by an alpha particle. By plotting the location of the flashes, it would be possible to determine how the path of the alpha particles through the atom was changed by the protons in the atom. The three paths shown in Figure (paths A, B, and C) are representative of those observed. Most of the alpha particles followed path A; they passed directly through the foil as though it were not there. Some were deflected slightly from their original path, as in path B; and an even smaller number bounced back from the foil as though they had hit a solid wall (path C).

Rutherford's experiment

Although you may be surprised that any alpha particles passed through the gold foil, Rutherford was not. He had expected that many would pass straight through (path A). He had also expected that, due to the presence in the atom of positively charged protons, some alpha particles would follow a slightly deflected path (path B). The fact that some alpha particles bounced back (path C) is what astounded Rutherford and his co-workers. Path C suggested that the particles had smashed into a region of dense mass and had bounced back. To use Rutherford's analogy, the possibility of such a bounce was as unlikely as a cannonball bouncing off a piece of tissue paper.

Results of the experiment

Careful consideration of the results and particularly of path C convinced Rutherford (and the scientific community) that an atom contains a very small, dense nucleus and a large amount of extranuclear space. According to Rutherford's theory, the nucleus of an atom contains all the mass of the atom and therefore all the protons. The protons give the nucleus a positive charge. Because like charges repel each other, positively charged alpha particles passing close to the nucleus are deflected (path B). The nucleus, containing all the protons and neutrons, is more massive than an alpha particle; therefore, an alpha particle striking the nucleus of a gold atom bounces back from the collision, as did those following path C.

Outside the nucleus, in the relatively enormous extranuclear space of the atom, are the tiny electrons. Because electrons are so small relative to the space they occupy, the extranuclear space of the atom is essentially empty. In Rutherford's experiment, alpha particles encountering this part of the atoms in the gold foil passed through the foil undeflected (path A).

If the nucleus contains virtually all the mass of the atom, it must be extremely dense. Its diameter is about 10-12 cm, about 1/10,000 that of the whole atom. Given this model, if the nucleus were the size of a marble, the atom with its extranuclear electrons would be 300 m in diameter. If a marble had the same density as the nucleus of an atom, it would weigh 3.3 X 1010 kg.

This model of the nucleus requires the introduction of a force other than those discussed earlier, one that will allow the protons, with their mutually repelling positive charges, to be packed close together in the nucleus, separated only by the uncharged neutrons. These nuclear forces seem to depend on interactions between protons and neutrons. Some are weak and some are very strong. Together they hold the nucleus together, but they are not yet understood.

The model of the atom based on Rutherford's work is, of course, no more than a model; we cannot see these subatomic particles or their arrangement within the atom. However, this model does give us a way of thinking about the atom that coincides with observations made about its properties. We can now determine not only what subatomic particles a particular atom contains but also whether or not they are in its nucleus. For example, an atom of carbon-12 contains 6 protons and 6 neutrons in its nucleus and 6 electrons outside the nucleus.

$^{12}_{6}C$

We have two distinct parts of an atom - the nucleus and the extranuclear space. The nucleus of an atom does not play any role in chemical reactions, but it does participate in radioactive reactions. (Such reactions are discussed later in this chapter.) The chemistry of an atom depends on its electrons - how many there are and how they are arranged in the extranuclear space.

The Mole

As we have observed, atoms are very small--too small to be weighed or counted individually. Nevertheless, we often need to know how many atoms (or molecules or electrons, and so on) a sample contains. To solve this dilemma, we use a counting unit called Avogadro's number, named after the Italian scientist Amedeo Avogadro (1776-1856):

Avogadro's number = 6.02 X 1023

Just as an amount of 12 is described by the term dozen, Avogadro's number is described by the term mole. A dozen eggs is 12 eggs; a mole of atoms is 6.02 X 1023 atoms. Avogadro's number can be used to count anything. You could have a mole of apples or a mole of Ping-Pong balls. You can get some idea of the magnitude of Avogadro's number by considering that a mole of Ping-Pong balls would cover the surface of the Earth with a layer approximately 60 miles thick. Avogadro's number is shown here to three significant figures, which is the degree of accuracy usually required in calculations. Actually, the number of items in a mole has been determined to six or more significant figures, the exact number depending on the method by which the number was determined.

One mole of any substance contains 6.02 X 1023 units of that substance. Equally important is the fact that one mole of a substance has a mass in grams numerically equal to the formula weight of that substance. Thus one mole of an element has a mass in grams equal to the atomic weight of that element and contains 6.02 X 1023 atoms of the element. For those elements that do not occur as single atoms - that is, the diatomic gases, sulfur, and phosphorus - it is important to be certain that you specify what you are talking about. One mole of atoms of oxygen has a mass of 16 g as 16 is the atomic weight of oxygen, and contains 6.02 X 1023 atoms of oxygen. One mole of oxygen gas, which has the formula O_2, has a mass of 32 g and contains 6.02 X 1023 molecules of oxygen but 12.04 X 1023 (2 X 6.02 X 1023) atoms, because each molecule of oxygen contains two oxygen atoms.

These definitions allow a new definition of atomic weight: The atomic weight of an element is the mass in grams of one mole of naturally occurring atoms of that element.

$$\text{Formula weight} = \frac{\text{grams}}{\text{mole}}$$

Using these relationships, we can calculate the number of atoms in a given mass of an element or the mass of a given number of atoms.