2 Atomic Structure

2.1 Introduction

At first glance, the periodic table may appear quite intimidating – a collection of seemingly unrelated numbers and symbols thrown together with little thought given to organization, continuity, or relevance. It may be comforting to know that the information contained in the periodic table is simply presented in a manner unfamiliar to most. A few basic skills are all that is required to appreciate the periodic table as a highly organized source of invaluable information.

2.2 Periodic Table

In 1869, a Russian chemist named Dmitri I. Mendeleev (1834–1907) discovered that elements exhibit a repeating pattern of properties when organized in order of increasing atomic mass. He called this observation the *periodic law*. Mendeleyev arranged the elements in different ways to determine if a relationship among the elements could be established. Frustrated and without success, he finally organized the elements into rows, beginning a new row with each repeating cycle, and the "primitive" periodic table was born. The German chemist Lothar Meyer also observed this "periodic" relationship around the same time, but as fate would have it, Mendeleyev is credited with the discovery because he was the first to organize the elements. The current form of the periodic table was later discovered by Henry Moseley who observed that elements organized by increasing atomic numbers created a more uniform arrangement.

The periodic table contains detailed information on all the known elements. The elements are represented by symbols that are translated into names. The symbols contain one-, two-, or three-letter designations that are case sensitive; the first letter is *always uppercase* with subsequent letters, if present, *always lowercase*. The names of the elements must be memorized along with their symbols. This task is somewhat simplified by the fact that many of the symbols share a common letter with the name. For example, N is the symbol for nitrogen and O is the symbol for oxygen. Unfortunately, this is not always the case, and some elements appear to have symbols that are completely unrelated to their names (Fig. 2.1). For example, Au is the symbol for gold and Fe is a symbol for iron, etc. In these cases, the symbol is derived from the element's name in Latin, German, or Greek; for example, Au is derived from the Latin word *aurum* meaning gold (see Appendix 1 for symbols and names of common elements). The elements are organized into vertical columns called groups and horizontal rows called periods. The group numbers appear at the top of each column in Roman numerals and range from I to VIII (one to eight). Elements in the same group (vertical column) have similar chemical and physical properties and a few groups are given characteristic names that you should become familiar with: Group IA elements are called alkali metals, group IIA are called alkaline earth metals, group VIIA the halogens, and group VIIIA the noble (or inert) gases. The A and B designations associated with the group numbers have no definitive meaning; they are simply used (in the US) to differentiate the main group elements (also called representative elements) from the transition metals. This practice can vary with table suppliers, particularly those from European countries. Nonetheless, the taller group columns located on each side of the table ("A" designations above) are called the main group elements and the middle groups ("B" designations above) are called the transition metals. The periods are numbered on the left side of the table from 1 to 7 downward, beginning with hydrogen.

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	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 H 1.01																	2 He 4.00
2	3 Li 6.94	4 Be 9.01											**5 B 10.81**	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	**14 Si 28.09**	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.71	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	**32 Ge 72.59**	**33 As 74.92**	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.91	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.87	48 Cd 112.4	49 In 114.8	50 Sn 118.7	**51 Sb 121.8**	**52 Te 127.6**	53 I 126.90	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.33	71 Lu* 175.0	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 196.97	80 Hg 200.59	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	**85 At (210)**	86 Rn (222)
7	87 Fr (223)	88 Ra (226.0)	103 Lr** (262)	104 Rf (261)	105 Db (262)	106 Sq (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)									

***Lanthanides**	57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (147)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
****Actinides**	89 Ac 227.0	90 Th 232.0	91 Pa (231)	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

Fig. 2.1 Periodic table of the elements (reprinted with permission from Dr. Thomas Basiri, Professor of Chemistry, Victor Valley College. Copyright 2008).

The periods extend left to right across the entire table. All periodic tables have a distinct line of separation running step-wise from boron/aluminum (B/Al) to polonium/astatine (Po/At). This line separates the metals to the left and the nonmetals to the right. The shaded elements bordering the line are called *metalloids* or *semi-metallics* because they possess both metallic (metal) and nonmetallic (nonmetal) properties. Metallic character decreases as you move left to right across a given period. Group IA and IIA elements are the most metallic and group VIIA and VIIIA are the most nonmetallic. A transition from metallic to nonmetallic character occurs through the group B elements and is the reason they are termed transition metals. The number above each symbol is called the atomic number and the number below each symbol is called the mass number. These numbers and their significance will be discussed below.

2.3 Atomic Structure

Elements are the *fundamental building blocks of matter*. Atoms are the *smallest, indivisible unit of an element that retains all chemical and physical properties* of the element. For example, a single atom of gold has the same physical and chemical properties as 10 tons of gold. Atoms are composed of three subatomic particles: protons, neutrons, and electrons. The protons and neutrons are located at the center of the atom in a region termed the nucleus. The mass number represents the *mass of the nucleus* or the total *mass of protons and neutrons*. The electrons are located in three-dimensional regions around the nucleus called *orbitals*. A large portion of any atom is empty space. The nucleus is surrounded by electrons in regions (orbitals) that are separated by great distances on a relative scale. Electrons located furthest from the nucleus, the *outermost electrons*, are called valence electrons and determine the chemical and physical properties of each element (Fig. 2.2).

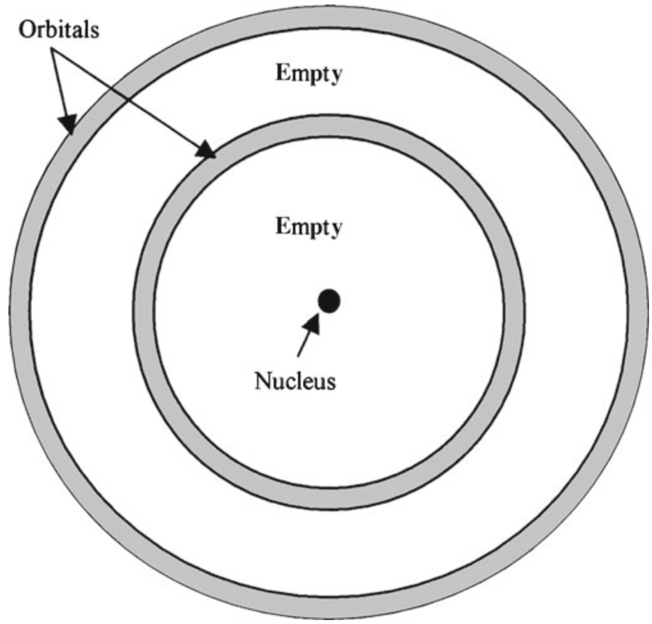


Fig. 2.2 The nucleus of subatomic particles is surrounded by electrons in orbital regions.

2.4 Subatomic Particles

Protons (p^+) are positively charged subatomic particles located in the nucleus. The *total number of protons* in the nucleus is given on the periodic table by the *atomic number* and *positively* identifies the element. A change in the number of protons changes the identity of the element; therefore, different elements *must* have different atomic numbers. A proton has a mass of one atomic mass unit (1 a.m.u = 1.66×10^{-24} g) and is comparable in size to a neutron.

Neutrons (n) are neutral subatomic particles also located in the nucleus. They carry no charge and therefore do not affect nuclear charge or the number of electrons in a neutral atom. Neutrons do contribute significantly to the mass of the atom (nuclear mass) because their mass is about the same as a proton. A change in the number of neutrons will change the mass number, but not the identity of the element. Atoms that contain the same number of protons but have different mass numbers are called *isotopes*. The number of neutrons contained in any nucleus is determined by subtracting the atomic number from the mass number.

Mass number = # of neutrons + # protons (atomic number)
Rearrange: # of neutrons = mass number – # protons (atomic number)

Electrons (e^-) are negatively charged subatomic particles located around the nucleus in predictable regions called *orbitals*. In a neutral atom, the number of electrons (negatives) is equal to the number of protons (positives). A change in the number of electrons in a neutral atom creates an ion, *an electrically charged atom*. Ions may carry a positive or negative charge depending on the number of electrons relative to the number of protons. The mass of an electron is approximately 2,000 times smaller than that of a proton or a neutron. As a result, electrons do not contribute significantly to the overall mass of the atom.

An atom is composed of two distinct regions – the nucleus and the region immediately surrounding the nucleus. The nucleus contains only protons and neutrons and is therefore positively charged. The region immediately surrounding the nucleus contains only electrons and is negatively charged. Protons and electrons are the only subatomic particles that are electrically charged. The charges are of equal magnitude despite the extreme disparity in mass, that is, an electron's charge will cancel a proton's charge. The overall charge on an atom is determined by comparing the number of protons to the number of electrons. If the number of electrons is *greater* than the number of protons (atomic number), the atom carries a *negative* charge; if the number of electrons is *less*, the atom carries a *positive* charge; if they are *equal*, the atom is *neutral*. The net charge on an atom in its standard state (natural conditions) is always zero (Fig. 2.3).

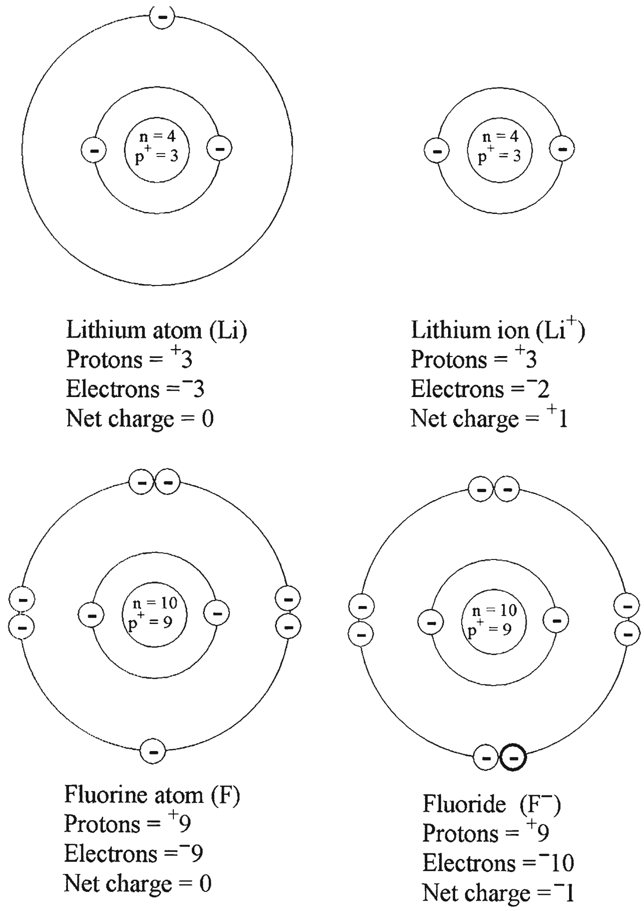


Fig. 2.3 Note how the net charge of protons and electrons determines if an atom's charge is neutral (0), positive (+), or negative (−).

2.5 The Arrangement of Electrons in an Atom

I grew up in a small town in upstate New York. On hot summer nights, our yard would fill with lightning bugs – interesting little flying creatures that periodically emit high-intensity light. When we were young, my sister and I would catch these elusive flying insects and put them in jars. We would punch out some small air holes in the lids, maybe throw in a piece of lettuce, and proudly display our new pets. Unfortunately, few ever saw the next night, a crushing reality to two small children. Imagine that you catch a single lightning bug and starve him for a few hours (incidentally, you did catch a male). You release your famished pet near a tree where you have graciously prepared a gourmet lightning bug dinner (you do, of course, know the diet of nocturnal insects!). You direct a camera at the food and take a time-lapsed photo over the remainder of the night (Fig. 2.4).

You would be very confident that, at any given time, there would be a high probability of finding the lightning bug in the region near the food. This statement is based on knowledge you possess, specifically, you know he is hungry. Your photo would most likely support your statement. You would see a region, or distribution, of light from the insect around the food. It is possible to find the lightning bug far from the food, but not likely given his current state of hunger. Electrons behave in a similar manner around the nucleus. It is not possible to know exactly where the electrons are; however, we can define regions where there is a high probability of finding them. Electrons are not randomly distributed around the nucleus; they are confined to specific energy levels called orbitals. At the subatomic level, it is not practical to use common units of length to measure distances, that is, it is not commonly stated that a specific electron may be 2 nm from the nucleus. Instead, we use energy to define distances. The electrons fill outward from the nucleus with the lowest energy, most stable electrons occupying regions close to the nucleus. The arrangement is very similar to an onion except great distances separate the

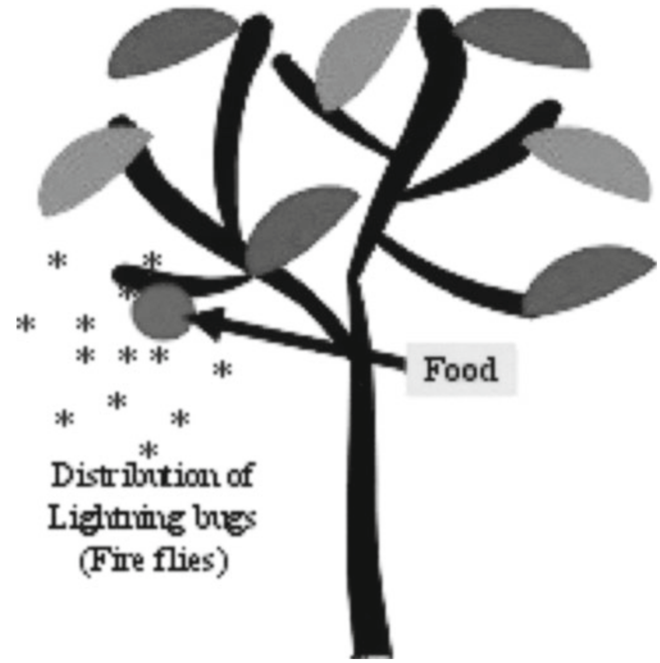


Fig. 2.4 The author's childhood experiment with lightning bugs shows a similarity with the relationship between electrons and the nucleus of an atom.

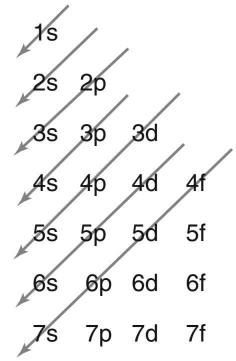
individual "peels." Thus, in comparing two electrons of different energy, it would be stated that electrons of higher energy are located further from the nucleus.

The energy of an electron is well defined and only certain energies are allowed (we call this quantized). If you have trouble with this statement, consider the musical tones created when you blow across the top of a bottle containing a fixed volume of water. It is not possible to create a full musical scale under these conditions. You hear only octaves or specific, allowable musical tones. The same can be said of an electron; it can have only specific, allowable energy values. This energy determines the location of the electron around the nucleus (remember, energy equates to distance; therefore, specific energy values translate to specific distances). The first levels of electron arrangement are "energy shells" called principal energy levels. They are given by the *period numbers* from the periodic table; for example, H and He represent *period 1* and have electrons in *principal energy level 1*, Li, Be, B, C, N, O, F, and Ne represent *period 2* and have electrons in *principal energy level 2*, and so forth. These regions (shells) increase in both size and energy as their distance from the nucleus increases. Principal energy level 1 is the lowest energy level (and smallest); therefore, electrons occupying this region are closest to the nucleus. Electrons occupying successively higher levels possess greater energy and are further from the nucleus. The *principal energy levels* are further *divided into sublevels*, or *orbitals*, designated *s*, *p*, *d*, and *f*. The orbitals are *regions around the nucleus where there is a high probability of finding an electron of specific energy*. Electrons occupy orbitals within principal energy levels and the energy of the electron determines which orbital it resides in and therefore its location around the nucleus. It may be helpful to compare electron arrangement to people staying at a hotel (I know…but bear with me!). First, hotels vary in size; some can have several floors with many rooms, while others have only a few floors with a small number of rooms (different atoms). Different people (the electrons) stay in different rooms (the orbitals) on different floors (the principal energy levels). Also, people are obligated to stay in only one room and generally cannot roam around and stay in any room they choose (specific energy of the electron).

2.6 Electron Configurations

Electron configurations illustrate the *arrangement of electrons around the nucleus of an atom*. The *aufbau principle* is used to construct electron configurations for ground-state (neutral) atoms and ions (*aufbau*: German for "build-up"). To determine the order that electrons fill around the nucleus, we must first construct the aufbau triangle (Fig. 2.5).

Fig. 2.5 An example of an aufbau triangle.



Draw the triangle as shown in Fig. 2.5; notice all rows contain the same number and all columns contain the same orbital (row three contains all 3's and the first column contains all *s*-orbitals). Draw parallel lines through the orbital designations as shown. Follow the arrows tail to head, beginning with 1s, and write the sequence: 1s2s2p3s3p4s3d…, etc. This is the sequence used to fill electrons around a nucleus. The electrons are located in orbitals (*s*, *p*, *d*, *f*) within principal energy levels (1, 2, 3, 4, etc.). To write an electron configuration, the number of electrons contained in the atom or ion must be calculated. Recall that protons and electrons are the only subatomic particles that carry a charge. In a neutral atom, the net charge is zero. This neutral state exists only when the number of electrons (negative charges) equals the number of protons (positive charges). Therefore, the number of electrons around the nucleus of a neutral atom is given by the atomic number. The following example illustrates how to write electron configurations. We will limit our discussion to *main group elements* only, that is, no *transition metals*.

Example: Write the electron configuration for Na.

First, determine the number of electrons in a neutral sodium atom. The periodic table gives an atomic number of 11 for Na. This means that there are 11 protons, or positive charges, in the nucleus of a sodium atom. Because no charge is written on the atom in our example, the number of electrons (negative charges) must also be 11.

Next, write a segment of the aufbau sequence.

$$1s2s2p3s3p4s3d$$

The number of electrons in each orbital is shown as a superscript attached to the orbital designation. The maximum number of electrons in an orbital: *s-orbital* is 2, *p-orbital* is 6, *d-orbital* is 10, and *f-orbital* is 14. The orbitals fill from *lowest to highest* energy and you *cannot* add electrons to higher levels until the preceding level is *full*. Revisiting our hotel analogy, you must fill the first floor before adding to the second; fill the second before adding to the third, etc. We start with the 1s orbital, principal energy level 1 containing a single s-orbital. We fill the orbital by placing a superscript 2 on the 1s designation (s-orbitals have a maximum occupancy of two electrons). Principal energy level 2 is filled next and contains a 2s and 2p orbital. The 2s orbital is filled in a manner similar to the 1s. The 2p orbital is filled using a superscript 6 attached (p-orbitals have a maximum occupancy of six electrons). We add the superscripts and find we have accounted for ten electrons. We have 11 electrons in total; so, the next orbital in our sequence (the 3s) will contain a single electron shown as a superscript of 1. The complete electron configuration for neutral Na is shown below.

$$1s^2 2s^2 2p^6 3s^1 3p4s3d$$

Our segment is too long so we simply erase the unused orbitals. If our segment was too short, we would add more orbitals to our sequence from the aufbau triangle.

The electron configuration for Na would be written as:

$$Na - 1s^2 2s^2 2p^6 3s^1$$

We would "read" this, 1s two, 2s two, 2p six, 3s one. The superscript on the last orbital depends on the number of electrons required to complete the configuration. It can be any number up to the maximum allowed in the orbital but can *never* exceed the maximum. Valence electrons are e*lectrons in the outermost principal energy level*. This may or may not be the last orbital written.

$$Na\left(11e^-\right) - 1s^2 2s^2 2p^6 3s^1$$

The outermost principal energy level containing electrons is level 3. Counting all electrons in level 3, we have 1 valence electron. In this case, the outermost level was the last one written.

$$O(8e^-)\text{–}1s^2 2s^2 2p^4$$

If we examine the electron configuration for oxygen, we see that the outermost principal energy level is 2. We have two orbitals in level 2 containing a total of 6 valence electrons, 2 in the 2s and 4 in the 2p (add the superscripts). In this case, the outermost level included the last two orbitals written. Care must be taken when determining the number of valence electrons; do not immediately jump to the last orbital and use its superscript. By chance, sometimes it is the last orbital, but sometimes it is not. Valence electrons occupy the highest principal energy level (level…level), not orbital (I think that I'll stop beating that horse now!).

The valence electrons determine the chemical and physical properties of the element. If two atoms were brought together during the course of a chemical reaction, their first point of contact would be the electrons in the outermost levels. If we know the electron configuration, and, therefore, the valence configuration, we can make predictions on properties and reactivity. Core electrons are located in levels below the valence electrons and generally do not influence reactivity.

The electron configurations for group I elements Li, Na, and K are shown below.

$$Li(3e^-)\text{–}1s^2 2s^1$$

$$Na(11e^-)\text{–}1s^2 2s^2 2p^6 3s^1$$

$$K(19e^-)\text{–}1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$$

If we look closely at the configurations, we see that each ends with s^1, a single electron in the outermost energy level. It is no coincidence that all the above elements are members of group I. The group numbers on the periodic table represent the number of valence electrons for each member of the group. Group I elements have 1 valence electron, group II elements have 2, group III elements have 3, etc.

Some members of group II:

$$Be(4e^-)\text{–}1s^2 2s^2$$

$$Mg(12e^-)\text{–}1s^2 2s^2 2p^6 3s^2$$

$$Ca(20e^-)\text{–}1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$$

Some members of group VII:

$$F(9e^-)\text{–}1s^2 2s^2 2p^5$$

$$Cl(17e^-)\text{–}1s^2 2s^2 2p^6 3s^2 3p^5$$

Electron configurations for ions are constructed in a similar manner except the charge must be considered in determining the total number of electrons. For ions, the number of electrons is calculated by subtracting the charge on the ion (with its sign) from the atomic number.

$$\# e^- (\text{for ions}) = \text{atomic number}(\#\ \text{of protons}) - (\text{charge})$$

Using the above relationship, we can calculate the total number of electrons contained in each of the following ions:

$$N^{3-}\ \text{N has atomic number 7}; \# e^- = 7 - ({}^{-}3) \text{ or } 10e^-$$

$$S^{2-}\ \text{S has atomic number 16}; \# e^- = 16 - ({}^{-}2) \text{ or } 18e^-$$

$$Br^-\ \text{Br has atomic number 35}; \# e^- = 35 - ({}^{-}1) \text{ or } 36e^-$$

$$Na^+\ \text{Na has atomic number 11}; \# e^- = 11 - ({}^{+}1) \text{ or } 10e^-$$

$$Ca^{2+}\ \text{Ca has atomic number 20}; \# e^- = 20 - ({}^{+}2) \text{ or } 18e^-$$

Notice that negative ions have more electrons than protons and positive ions have fewer electrons than protons. Ions are created by changing the number of electrons relative to the number of protons. It is worth noting that moving around electrons is not rocket science, if you have ever rubbed a balloon on your head and stuck it on a wall you have accomplished this miraculous feat. Below are the electron configurations for N^{3-}, O^{2-}, F^-, Ne, Na^+, and Mg^{2+}.

$$N^{3-}\left(10e^-\right)-1s^2\,2s^2\,2p^6$$
$$O^{2-}\left(10e^-\right)-1s^2\,2s^2\,2p^6$$
$$F^-\left(10e^-\right)-1s^2\,2s^2\,2p^6$$
$$Ne\left(10e^-\right)-1s^2\,2s^2\,2p^6$$
$$Na^+\left(10e^-\right)-1s^2\,2s^2\,2p^6$$
$$Mg^{2+}\left(10e^-\right)-1s^2\,2s^2\,2p^6$$

The above configurations are identical and each contains a total of 10 electrons with eight valence electrons (remember our dead horse; valence electrons occupy the highest principal energy level, which is 2 in this case). Although the configurations are identical, the charges on the ions vary and only one atom in our group is neutral. Neon is a "noble gas" or "inert gas," names given to all group VIIIA elements. Chemical reactivity is a stability-driven process – if products are more stable than reactants (starting material), the reaction occurs. The noble gases are extremely stable and show very little reactivity because of filled outer shell configurations. In the above examples, principal energy level 2 (the outer level or shell) is full; principal energy level 2 has only s and p orbitals. The addition of a single extra electron to any of the above configurations would require occupancy in the next higher energy shell, specifically the 3s orbital. A filled outer shell configuration is achieved when all orbitals in the outermost level are full. Generally, this is achieved with eight valence electrons. *The tendency of atoms to gain or lose electrons to obtain electron configurations similar to group VIII elements is called the* octet rule. The configurations above, all satisfy the octet rule and represent the elements in their most stable forms. The charges on the atoms result from a gain or loss of electrons in order to achieve a configuration identical to Ne. There are exceptions to the octet rule. For example, helium (He) is a group VIIIA element that has an electron configuration of $1s^2$. Principal energy level one contains only the 1s-orbital and requires only two electrons to fill. Helium satisfies this condition and has therefore achieved its octet.

There is a difference between the ground state of an element and its most stable state. The ground state is how the atom is most often found in nature and occurs when the atom is neutral. The most stable state is when the atom has achieved an octet, or filled outer shell configuration. This will generally require the atom to carry an overall net charge as a result of gaining or losing electrons. The group numbers provide us with important information on both states. We know that the number of valence electrons for a particular atom is given by the group number. However, what may have escaped our attention is that the group number gives the valence electrons for neutral atoms only; group I elements have 1 valence electron, group II have two, etc. Let us revisit our configurations above for N^{3-}, O^{2-}, F^-, Ne, Na^+, and Mg^{2+}; all contain eight valence electrons regardless of their group number, but none are neutral except Ne (you guessed it, a group VIIIA element). Consider the following examples, paying close attention to the differences in electron configuration and how this relates to charge in the most stable state:

Natural state of sodium – Na ($11e^-$) – $1s^22s^22p^63s^1$
Most stable state of sodium – Na^+ ($10e^-$) – $1s^22s^22p^6$
Natural state of magnesium – Mg ($12e^-$) – $1s^22s^22p^63s^2$
Most stable state of magnesium – Mg^{2+} ($10e^-$) – $1s^22s^22p^6$
Natural state of aluminum – Al ($13e^-$) – $1s^22s^22p^63s^23p^1$
Most stable state of aluminum – Al^{3+} ($10e^-$) – $1s^22s^22p^6$
Natural state of nitrogen – N ($7e^-$) – $1s^22s^22p^3$
Most stable state of nitrogen – N^{3-} ($10e^-$) – $1s^22s^22p^6$
Natural state of atomic oxygen – O ($8e^-$) – $1s^22s^22p^4$
Most stable state of atomic oxygen – O^{2-} ($10e^-$) – $1s^22s^22p^6$
Natural state of atomic fluorine – F ($9e^-$) – $1s^22s^22p^5$
Most stable state of atomic fluorine – F^- ($10e^-$) – $1s^22s^22p^6$

Relating the group number to the octet rule can provide information on the most stable state of group members. Group I elements have one valence electron and can obtain an octet (eight electrons in the outermost principal energy level) in two ways: lose one electron and take a $^+1$ charge or gain seven electrons and take $^-7$ charge. The more favorable choice would be to lose one electron and become $^+1$. Groups IA, IIA, and IIIA lose electrons to achieve their octets and take charges of $^+1$, $^+2$, and $^+3$, respectively, in their most stable forms. Group IV is unusual and will not be discussed at this point. Groups VA, VIA, and VIIA will gain electrons to complete their octets, taking charges of $^-3$, $^-2$, and $^-1$, respectively, in their most stable forms. We can summarize these observations by stating: *to satisfy the octet rule and achieve stable electron configurations, metals must lose electrons and nonmetals must gain electrons*.

Table 2.1 Orbitals within each principal energy level and the maximum number of electrons contained in each

Orbital		Maximum number of electrons	
s		2	
p		6	
d		10	
f		14	
Principal energy level	Number of orbitals	Orbital designation	Maximum number of electrons
1	1	1s	2
Maximum number of electrons in principal energy level 1 is 2			
2	2	2s	2
		2p	6
Maximum number of electrons in principal energy level 2 is 8			
3	3	3s	2
		3p	6
		3d	10
Maximum number of electrons in principal energy level 3 is 18			
4	4	4s	2
		4p	6
		4d	10
		4f	14

Table 2.2 Electron distribution in atomic orbitals

Principal energy level (n)	Formula $2n^2$	Maximum number of electrons
1	$2\times(1^2)$	2
2	$2\times(2^2)$	8
3	$2\times(3^2)$	18
4	$2\times(4^2)$	32

Table 2.1 shows the orbitals within each principal energy level and the maximum number of electrons contained in each. Notice that not all principal energy levels contain all orbitals.

The maximum number of electrons contained in a principal energy level can be calculated using the formula $2n^2$, where n is the principal energy level (period number from the periodic table) (Table 2.2).

2.7 Periodic Trends: Understanding the Periodic Table

The periodic table is an arrangement of the elements based on similarities in atomic properties. It can therefore be used to predict the chemical and physical properties of elements. Electronegativity is a measure of an atom's *desire for electrons* or the ability of an atom to draw electrons toward it. Fluorine (F) is the most electronegative element on the periodic table and, in general, the periodic trend is that the closer an element is to fluorine, the greater is its electronegativity. If asked which element, Br or Ca, has the greater ability to draw electrons, you would respond Br because it is closer to fluorine. The nonmetals are grouped near fluorine on the periodic table and must gain electrons to achieve octets; as a result, they have high electronegativities. Metallic elements, particularly groups IA and IIA, are not near fluorine on the periodic table. They must lose electrons to achieve octets and therefore have low electronegativities. The importance of electronegativity in chemical bonding cannot be overstated and a good rule of thumb is "the closer an element is to fluorine on the periodic table, the greater is its electronegativity"

Theoretically, the atomic radius of an atom is the distance from the center of the nucleus to the outer boundary of the atom. However, regions containing the valence electrons (outermost) do not have distinct boundaries. Not to worry; when two atoms of the same element are bound together, the centers of their nuclei are separated by a measurable distance. Therefore, the atomic radius of an atom is *defined as half the distance between the centers of two bonded atoms of the same element*; for example, the atomic radius of a single hydrogen atom is equal to half the distance from the centers of two bonded hydrogen atoms. Atomic radii increase down a given group of elements and, in general, decrease left to right across a period (Fig. 2.6).

Atomic radius decreases

Atomic radius decreases

Electronegativity increases

	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 H 1.01																	2 He 4.00
2	3 Li 6.94	4 Be 9.01											**5 B 10.81**	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	**14 Si 28.09**	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.71	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	**32 Ge 72.59**	**33 As 74.92**	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.91	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.87	48 Cd 112.4	49 In 114.8	50 Sn 118.7	**51 Sb 121.8**	**52 Te 127.6**	53 I 126.90	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.33	57 La* 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 196.97	80 Hg 200.59	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	**84 Po (209)**	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226.0)	89 Ac** (227)	104 Rf (261)	105 Db (262)	106 Sq (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)									

Electronegativity increases

Fig. 2.6 Understanding periodic trends.

2.8 Isotopes

Atoms that contain the same number of protons but have different mass numbers are called isotopes. Isotopes of an element differ only in the number of neutrons contained in the nucleus. Typically, atomic nuclei are most stable when they contain a certain number of protons and neutrons. The addition of neutrons to the nucleus increases the mass of the atom and creates instability. For example, the most abundant form of hydrogen contains a single proton in the nucleus. The addition of a neutron to a hydrogen nucleus creates an isotope of hydrogen called deuterium. Deuterium is a heavier and more energetic form of hydrogen, and is therefore less stable. The addition of a second neutron creates a third isotope of hydrogen called tritium, the most unstable and active form of hydrogen. The instability of isotopes is a direct result of increased nuclear mass and is detected through a release of energy called nuclear radiation (Fig. 2.7).

2.9 Radioactivity

Dig a ditch in 100° or take a nap on the couch; well, there is a tough choice. The relationship between energy and stability is a common thread that unifies most areas of science (and based on your response to my question, it appears that it extends into our daily lives as well). High energy translates to instability and there is always a natural tendency toward lowest energy and greatest stability (the nap on the couch). This is inescapable; however, I would not try this argument the next time your asked to do yard work, it does not work, believe me. The response of an atom to high energy is not much different from our own. It will not remain unstable indefinitely; eventually, the nucleus will emit energy in an effort to regain stability (its version of a nap). The *spontaneous emission of high-energy nuclear radiation from an unstable nucleus* is termed radioactivity (or radioactive decay). Atoms that exhibit this property are said to be radioactive and most elements with an atomic number of 90 or greater have radioactive isotopes. Early experiments identified three types of nuclear radiation: alpha (α), beta (β), and gamma (γ) rays. A sample of radioactive material is placed between the positive and negative poles of a magnet and emitted radiation is detected using a piece of X-ray film placed at the top of the apparatus (Fig. 2.8).

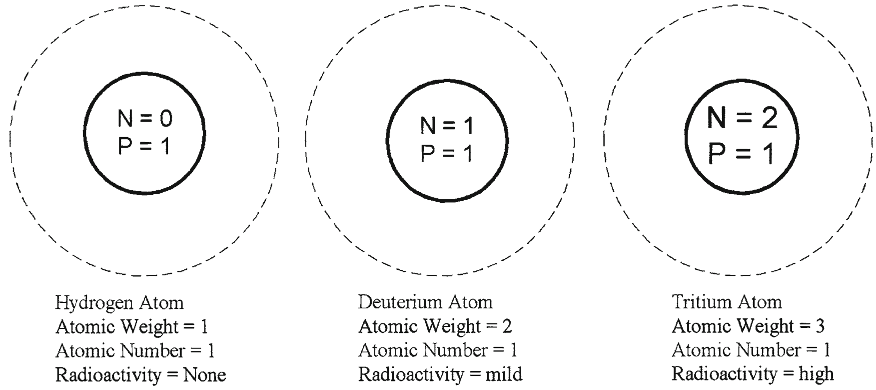


Fig. 2.7 As the nuclear mass of an element increases, so does the isotopes instability.

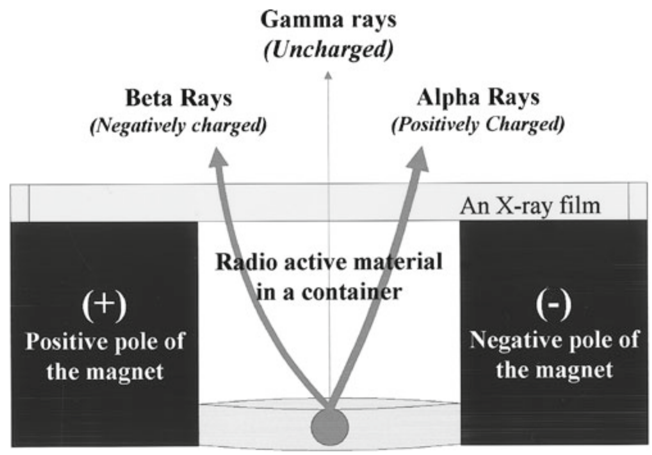


Fig. 2.8 A magnet is used to measure the emission of radioactivity from sample material.

Three spots were observed at different locations on the film. Two spots were deflected toward opposite poles of the magnet, whereas the third passed straight through, apparently unaffected. This implies that two of the particles are electrically charged and the third is neutral. Alpha (α) rays are positively charged particles that are deflected toward the negative pole and beta (β) rays are negatively charged particles deflected toward the positive pole. Gamma (γ) rays have no detectable charge (or mass) and therefore passed straight through.

2.10 Types of Radioactive Decay

The release of a helium nucleus (He^{2+}) during radioactive decay is called α-decay (alpha decay). This type of decay is a low-energy emission of positively charged particles. A thin sheet of paper will provide adequate protection against this type of radiation. The release of electrons (e^-) during radioactive decay is called β-decay (beta decay). This type of decay produces negatively charged particles of medium energy. A few hundred sheets of paper are required to provide adequate protection against this type of radiation. The release of electromagnetic radiation during radioactive decay is called γ-decay (gamma decay). This type of decay produces high-energy, neutral radiation capable of penetrating a 1-inch-thick wall of lead. This is the most dangerous and destructive form of radioactive decay.

2.11 Nuclear Radiation: Forensic Applications

Radioactive isotopes will lose intensity (gain stability) over time because of α-, β-, or γ-decay. The amount of time required for radioactive intensity to decrease by half is called the half-life. Carbon-14 is a radioactive isotope of carbon with a half-life of 5,720 years. A 100-g sample of radioactive carbon-14 will contain 50 g of active carbon-14 in 5,720 years, 25 g after an additional 5,720 years, and so on. Half-lives can range from fractions of a second to millions of years, depending on the isotope.

Table 2.3 The atomic number, atomic mass, and molar mass of selected elements

Element	Atomic number	Mass of 1 Atom (atomic mass)	Mass of one mole (molar mass)
H (hydrogen)	1	1.01 a.m.u.	1.01 g
He (helium)	2	4.00 a.m.u.	4.00 g
Li (lithium)	3	6.94 a.m.u	6.94 g
C (carbon)	6	12.01 a.m.u.	12.01 g
N (nitrogen)	7	14.01 a.m.u.	14.01 g
O (oxygen)	8	16.00 a.m.u.	16.00 g
F (fluorine)	9	19.00 a.m.u.	19.00 g
Na (sodium)	11	22.99 a.m.u.	22.99 g
P (phosphorous)	15	30.97 a.m.u.	30.97 g
Cl (chlorine)	17	35.45 a.m.u.	35.45 g
I (iodine)	53	126.90 a.m.u.	126.90 g

Forensic anthropologists use this information to determine the age of ancient artifacts, mummies, bones, and other material. Radioactive dating is a common technique accepted worldwide.

2.12 The Mole and Molar Mass

The atomic mass of carbon from the periodic table is 12.01, but 12.01 of "what"? Curiously, no "mass" units are given on the periodic table with "mass numbers." The reason is that mass numbers can have two equally important units: *atomic mass units* (a.m.u.) or *grams*. The preferential inclusion on the table of one unit over the other would undoubtedly spark a never-ending debate, dividing educators and authors worldwide. To avoid this debacle, and the certain demise of the modern world, no units are given; after all, the last thing we need is another source of debate. I momentarily digress, let us return to carbon: 12.01 amu's of carbon represents the mass of *one carbon atom*, 12.01 g of carbon represents the mass of 6.02×10^{23} *atoms* of carbon. The mass number in amu's of any element represents the mass of one atom of the element, whereas the mass number in grams represents the mass of 6.02×10^{23} atoms of the element. The mass "numbers" are the same; it is the units that distinguish the difference. If you were asked how many pencils are in a dozen pencils, you would reply 12. We associate the word "dozen" with the number "12" and define a dozen as anything that contains 12 "things." The same is true of a mole, an extremely important quantity used in chemistry. A mole is defined as anything that contains 6.02×10^{23} particles or "things." We associate the number 6.02×10^{23} with the word "mole." We can simplify our example above by stating: *the atomic mass of any element, in grams, contains* 6.02×10^{23} *atoms of the element, or one mole of the element, and is called the molar mass*. The quantity that defines a mole, 6.02×10^{23}, is called Avogadro's number in honor of its founder, the nineteenth-century Italian scientist Amadeo Avogadro.

2.13 Elements of Forensic Interest

See Table 2.3 for the elements of forensic interest.

2.14 Questions

1. Write the names of the elements represented by the following symbols:
   (a) I
   (b) P
   (c) Na

2. Write the symbols for the following elements:
   (a) Potassium
   (b) Nickel
   (c) Manganese
   (d) Magnesium
3. Name the three types of subatomic particles and give their location in the atom.
4. Provide the mass that contains:
   (a) One atom of carbon
   (b) One mole of magnesium
   (c) 6.02×10^{23} atoms of Li
   (d) 3.01×10^{23} atoms of Ca
5. Please explain to the members of the jury how two atoms of the same element can have different mass numbers.
6. Define radioactivity and the three types of nuclear radiation.
7. Cite a few examples of the application of radioactive decay to forensic investigation.
8. Explain the aufbau principle.
9. Give the maximum number of electrons in:
   (a) Principal energy level 2
   (b) A *p*-orbital
   (c) Principal energy level 4
   (d) The 4f-orbital
   (e) The 1s-orbital
10. Briefly explain to the members of the jury the difference between a neutral atom and an ion. How are ions formed?
11. Write the electron configuration for each of the following:
   (a) Na
   (b) F^-
   (c) Mg^{2+}
   (d) Li^+
   (e) Ar
12. Explain why the electron configurations for N^{3-}, O^{2-}, F^-, Ne, Na^+, and Mg^{2+} are identical.
13. What information does the group numbers of the periodic table give?
14. Describe the difference between the natural state of an atom and its most stable state.
15. Describe the periodic trends of electronegativity and atomic radius.

Suggested Reading

Jones, L.; Atkins, P. *Chemistry: Molecules, Matter, and Change*, 4th ed.; W.H. Freeman and Company: New York, 2002; pp 298–299, pp 959–964.

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