

Chapter 8. The Periodic Table

8.1. Chemical Families

There is a natural tendency to search for order in the Universe. The “secrets” of Nature that scientists attempt to discover are relationships *between* things; causes produce effects, objects obey laws, behavior should be predictable.¹ Chemists have always sought for relationships among material substances. Soon after John Dalton showed how to determine the relative atomic masses of atoms of the elements, chemists were seeking for relationships between atomic mass and chemical and physical properties. Numerous Nineteenth Century chemists speculated about *families* of elements having similar chemical properties. For example, lithium, sodium and potassium form a family of highly reactive metallic elements; their speed of reaction with water, for example, increases proportionately in the order listed, as does their relative mass, suggesting that the properties of sodium are averages of those of lithium and potassium. Relationships between chemical properties and mass could only be sketchy at first because so few elements were known and properties were not well established. By the middle of the Nineteenth Century, however, enough information was available for a general pattern to emerge. One effort stood above the others in finding a grand classification scheme for all the known elements. In 1871 Dimitrii Mendeleev² published a chart of the 63 elements known to his time, arranged in rows with atomic mass increasing regularly along each row. It could be constructed by first listing all the elements along a horizontal line according to increasing atomic mass, then breaking the line into segments and rearranging the segments into a series of parallel rows. The rows were placed under each other such that families of elements having similar properties lined up in columns, much the same way

¹ Regarding predictability, Einstein once expressed his aversion to the statistical interpretation of wave mechanics by stating, “God does not play dice!”

² Dimitrii Ivanovitch Mendeleev (Siberian, 1834-1907) Nobel Prize in chemistry in 1903. After her school-teacher husband’s death, Dimitrii’s intrepid mother took her youngest of seventeen children to St. Petersburg to continue his education in science. The periodic chemical arrangement came to him in a dream in which he saw cards with the properties of the elements he had prepared arrange themselves into columns of families.

calendars display a string of a year of days in rows of weeks with columns of days of the week. Mendelyev's chemical "calendar" required occasional *gaps* where the mass took larger jumps than usual, suggesting the possibility of yet undiscovered elements. The resulting chart of the elements is called the **chemical periodic table**, or **periodic chart** because of the periodic recurring resemblances along the horizontal original line of increasing atomic mass. A modern periodic table is given inside the front cover of this book. Unlike Mendeleev's original table, which had only a few small gaps for undiscovered elements, the modern form has large gaps, which will be shown below to be associated with aufbau electronic structure ordering.

Mendeleev's table was distinctive because he recognized that families of elements with similar properties could be arranged into columns only by leaving vacancies in certain places and also by occasionally reversing pairs in others. Mendeleev gave tentative names to the missing elements, and predicted their properties from analogy with properties of their known relatives. These predictions proved prophetic as new elements were discovered. The inversions in mass order he attributed to inaccurate mass determinations, which proved to be the case in some instances, but remained a mystery for others until the periodic table was put on a theoretical footing in the Twentieth Century based on subatomic particles.

Example 8.1 Averaging the atomic masses of calcium and titanium, 40 and 48, respectively, Mendeleev predicted the mass of the undiscovered element between them to be 44. The atomic mass of the new element (named scandium because it was discovered in Scandinavia) turned out to be 45.

8.2. Family Relationships

The arrangement of the elements into 20 families shown on the inside front cover (18 vertical families above and 2 horizontal families below) summarizes a variety of trends in chemical and physical properties of the elements. In broad terms the "long form" of the periodic table³ separates the *metallic elements* on the *left* from the *non-metals* on the *right*.

³ Only one of dozens of different proposed arrangements, the long form of the periodic table is a compromise between Mendeleev's short form, which integrated the transition metals into blocks at the ends, and the "very long" form, which makes room for the fourteen columns of the rare earth elements (lanthanides and actinides). The long form's chief advantage is that it conforms nicely to the dimensions of rectangular paper in textbooks. When textbooks are replaced by graphics display devices, there may be a revival of three-dimensional forms, which display periodicity in a cyclic spiral fashion.

Families of elements fall into larger groups of families consisting of chemically *active metals* (first two columns on the left), *representative elements* (first two columns on the left plus last six columns on the right) where the heads of each family are representative of the remaining members, *transition elements* bridging the transition between the left and right representative elements, and the *lanthanide* and *actinide rows* separating the active metals from the transition metals. Historically, certain families of elements were grouped together according to chemical properties. The first column of active metals on the left is called the *alkali metal* family, the second *alkaline earths*, the lanthanides and actinides together are called *rare earths*, and the last three representative families are called *calcogens*, *halogens* and *noble gases*. (The remaining representative families are named for their top members, boron, nitrogen and carbon.) Because hydrogen and helium are electronically similar to the alkali and alkaline metals, respectively, but are chemically similar to the halogens and noble gases, respectively, they may be placed in either group. It is useful to know that although there are common unstable isotopes of most of the elements, there are *no stable isotopes of elements beyond bismuth* (atomic number 83), and that *no elements beyond uranium* (atomic number 92) are found in nature (they are man made in nuclear reactions).⁴

The columns of the periodic table summarize trends in similar chemical properties such as reactivity (greater at the extremes) and combining power, or **valence**, which indicates the number of atoms a given atom combines with. The maximum valence for the representative elements is the family, “group”, or column number of the element for the first four columns, and (8 – the column number) for the remaining four columns, according to the column numbering scheme given on the inside front cover.

Chemists frequently use the heuristic of *reasoning by analogy*. It is based on the expectation that similar systems behave similarly. This is the reasoning Newton used to deduce the behavior of the moon from the behavior of an apple; both are attracted to the earth by similar gravitational forces. Knowing that the chemical elements fall into families with similar chemical properties suggests a powerful heuristic based on analogous chemical properties. The chemical periodic heuristic has been justified through the understanding of electronic structure of atoms developed in the Twentieth Century.

⁴ While they may be generated in natural stellar processes, their lifetimes are so short they would decay into other elements before they can be detected.

The Periodic Law Heuristic

Purpose: To determine the properties of a given substance (element or compound).

Procedure:

1. Locate the families to which the elements of the substance belong in the periodic table.
2. Recall the properties of substances with elements from the same families and infer similar properties to the given substance.

While step 2 requires some degree of chemical knowledge or experience, it does compress the amount of knowledge needed to treat a large variety of analogous compounds. There are 50 representative elements, which can combine in 2,500 possible ways to form binary molecules. However, there are only eight families of representative elements. Thus knowledge of the properties of only 8 representative binary molecules allows one to predict the properties of the remaining $(42-8) \times (42-8) = 1,156$ possible pairs. The advantage increases with the number of atoms in the molecules ($(42 - 8)^n$ vs 8^n , for $n = 2, 3$, etc. atoms).

Example 8.2 Lithium metal reacts with water to produce hydrogen gas and a basic solution. What would be expected for other alkali metals?

1. Lithium stands at the head of the alkali metal family column.
2. Sodium, potassium, rubidium and cesium should also react with water to produce hydrogen gas and a basic solution. The truth of the prediction is demonstrated by a simple experiment.

Example 8.3 What can be inferred from the fact that gold is a rare, pure substance.

1. Gold is a soft metal which belongs to the family that includes copper and silver.
2. Copper and silver occur as uncombined elements in nature, too. The rarity, inertness and malleability of these elements is reason to use them for coins. This family is called the “coinage metal” family.

Example 8.4 What can be inferred from the knowledge that “milk of magnesia” (an alchemical name for the compound magnesium hydroxide) neutralizes acids.

1. Magnesium is among the representative metals (left two columns of the periodic table).
2. The inference is that the other representative metals neutralize acids too. Substances that neutralize acids are called *bases*, several common members of which contain alkali metal and alkaline earth elements.

Example 8.5 Discuss some properties of the halogens.

1. Common chemical knowledge:

Iodine, a member of the halogen family, is an antiseptic.

Table salt contains chlorine and is a spice and preservative.

Industrial muriatic acid (hydrogen chloride dissolved in water) is a strong acid (largely ionized in water solution).

2. Halogens should display destructive biological activity. Chemists are familiar with the fact that chlorine and bromine attack the mucus membranes of the nose, throat and lungs.

Other metal compounds of halogens are known to taste “salty.” Lithium chloride is used as a salt “substitute” for low sodium diets.

Other halogen acids, such as bromic and iodic acid are also strong acids.

Analogous relationships are not always correct. In the last example, it would be a mistake to classify hydrofluoric acid as a strong acid, in the usual sense of being strongly ionized, the usual characterization of strong acids (although it does attack glass). Nevertheless, enough similarities exist to make the periodic table a reasonably good predictive tool.

8.3. Periodic Configurations

Lets explore the electronic configurations of a few neutral elements having low atomic numbers (numbers of protons). We will use the Aufbau Electronic Configuration Heuristic of Chapter 7 as our guide. Beginning with hydrogen (atomic number 1), the ground state electronic configuration is $1s^1$, indicating one electron in the 1s subshell. Helium (atomic number 2) has the electronic configuration $1s^2$, representing two electrons with two unique configurations $\{1,0,0,-\frac{1}{2}\}$ and $\{1,0,0,+\frac{1}{2}\}$.⁵ The element with atomic number three, lithium, has

⁵ Cf. Section 7.6 on Atomic Electronic Configurations.

three electrons, the third of which must find a new shell according to the exclusion principle, as the first shell is fully occupied by two of the electrons ($2n^2 = 2$ for $n = 1$). Thus the electronic configuration of lithium is $1s^2 2s^1$. Electronically, the outer-subshell electron configuration (ns^1) is similar to that of hydrogen ($1s^1$). Continuing in this way, the next (ns^1) outer-subshell configuration occurs at atomic number 10 (sodium, $3s^1$), and the next at atomic number 19 (potassium with $4s^1$). It happens that lithium, sodium and potassium are highly reactive metals, having similar chemical and physical properties. (Hydrogen is a gas, but can exist in a metallic state under extremely high pressures, such as found in the interior of the sun and Jupiter.)

Exploring other elements, we discover more similarities in harmony with the generalization that *similar outer-subshell electronic configurations correlate with similar chemical and physical behavior*.

Figure 8.1 shows how the first 118 neutral elements would be grouped into columns having similar outer-subshell electronic configurations, if arranged according to increasing atomic number in horizontal rows. The gaps are the result of the aufbau ordering of electronic configurations and the restrictions on the ranges of allowed quantum numbers.

1s			
2s			2p
3s			3p
4s		3d	4p
5s		4d	5p
6s	4f	5d	6p
7s	5f	6d	7p

Fig. 8.1 Block Form of the Periodic Table Showing Subshell Structure

This arrangement is called the *electronic structure periodic table* due to the periodically recurring properties with increasing atomic number. Each ns block ($n = 1, 2$, etc.) represents subshell occupancies of s^1 and s^2 , each np block represents subshell occupancies of $p^1 \cdots p^6$,

each (n-1)d block represents subshell occupancies of $d^1 \cdots d^{10}$, and each (n-2)f block represents subshell occupancies of $f^1 \cdots f^{14}$, corresponding to the possible occupancies of subshells allowed by the results of wave mechanics.

Accurate calculations of electronic structure show some exceptions to the aufbau heuristic for ordering energies of ground state atomic electronic configurations. Slight variations in subshell occupancies occasionally occur among the heavier elements (beyond atomic number 23), where the subshell energies become nearly degenerate (of the same value). A footnote to the periodicity of electronic configurations is the rule that *half-filled and filled subshells are particularly stable*. Thus according to the strict aufbau ordering, copper would have a ground state electronic configuration of $[\text{Ar}]4s^23d^9$, whereas in fact this is an observed excited state lying slightly above the observed ground state configuration of $[\text{Ar}]4s^13d^{10}$, which is favored by a combined filled subshell and half-filled subshell. Another notable exception occurs at the beginning of the f subshell occupancies, which are preempted by the next d subshell. This results in the positioning of lanthanum and actinium in the scandium family in traditional forms of the chemical periodic table, although a strict observance of the aufbau ordering would place them with the other inner transition series. (Cf. inside front cover.)

Ionization is removal of electrons to form *cations* and removal of electrons should follow the aufbau filling in reverse. However, experimental and theoretical evidence shows that *outer shell electrons are removed before (with less energy than) inner shell electrons, and within a shell, higher quantum number orbitals are removed more easily than lower quantum number orbitals*. Thus copper forms two common cations, Cu^+ by removal of the $4s^1$ outer shell electron from the ground state, and Cu^{2+} by removal of the $4s^2$ electrons from the first excited state.

The chemical periodic table coupled with the electronic configuration periodic table of Fig. 8.1 permits a quick identification of the electronic configurations of the elements, according to the orbital approximation and the aufbau heuristic (not including occasional occupation reversals). Since each row of Fig. 8.1 corresponds to filling subshells, the row number of the table reflects the principle quantum number n of the subshell being filled; for s and p subshells the row number equals n , for d subshells the row number is $n-1$, and for f subshells the row number is $n-2$.

Noble gas elements have filled p subshells of electrons, and are chemically stable, or non-reactive. In other elements noble gas configurations correspond to inner subshell **core**

electrons and may be summarized by noting the noble gas atom to which they correspond. The outer unfilled subshell, or **valence electrons**, which are responsible for the majority of the chemical properties of the atoms may then be appended to the core configuration to complete the electronic configuration.

The Periodic Table Electronic Configuration Algorithm

Purpose: To determine the ground state electronic configuration of an atom or ion from the chemical periodic table.

Procedure: The periodic table is arranged in *blocks* of s, p, d and f subshells, having 2, 6, 10 and 14 columns each, respectively. To determine the electronic configuration of an atom:

1. If the atom has charge (ionized), use the equivalent electronic configuration for the neutral atom to the left (for positive charge) or right (for negative charge) of the given element, shifted by a number equal to the magnitude of the charge.
2. Locate in the periodic table the noble gas with the nearest preceding atomic number to the neutral element from step 1 and summarize its electronic configuration with the chemical symbol of the noble gas in square brackets, []. This accounts for as many electrons of the element as the noble gas has.
3. Locate the row number, r , and block symbol, l (for s, p, d, or f), of the given atom in the periodic table. Scan the row of the element and append to the filled subshell configuration of step 2 additional configurations of the form nl^{occ} for each block preceding the block of the element, using $n = r$ for s and p block electrons, $n = r - 1$ for d block electrons, and $n = r - 2$ for f block electrons. l stands for the l quantum number block symbol (s, p, d or f) and occ equals 2 for s, 6 for p, 10 for d and 14 for f blocks.
4. Locate the column number, c , of the block in which the neutral element resides. To the configuration of step 3, append the electronic configuration of the remaining electrons as nl^c , where n and l are as given in step 3.

Example 8.6 Write the electronic configuration of neutral lead in the ground state.

1. The neutral atom has zero charge, so find the position of lead (Pb) in the periodic table.
2. The nearest noble gas preceding lead (Pb) is xenon. Its symbol is Xe. The symbol [Xe] represents the electronic configuration of the 54 lowest energy electrons in lead.
3. Pb is in the sixth row and in the p block. Add $6s^25d^{10}4f^{14}$ to the electronic configuration of Xe.
4. Pb is in the $n = 6$ row and second column from the left in the p block. Add $6p^2$ to the electronic configuration of step 2. The final configuration is $[\text{Xe}]6s^25d^{10}4f^{14}6p^2$.

This example points out a potential difficulty of displaying the periodic table in the long form, rather than the very long form of Fig. 8.1. If one doesn't notice the gap between elements 57 and 72 in the long form of the periodic table, one is likely to overlook the 14 4f electrons.

8.4. Periodic Trends

Fairly regular trends of behavior exist along the rows and columns of the periodic table. For example, among the columns of the representative elements which contain both metals and non-metals, metallic behavior increases downward. Thus carbon is classified as a non-metal⁶, silicon and germanium as *metalloids* (or *semi-conductors*), and tin and lead exhibit the properties of hardness and conduction common to the metals. Among the alkali metals and alkaline earth columns, melting point decreases downward and chemical reactivity increases.

A number of properties of atoms depend essentially on *size*, measured in terms of atomic volume or radius. A remarkable near cancellation of decreasing size with increasing

⁶ Elemental carbon occurs in a number of forms, depending on the arrangement and bonding of its atoms, variously known as *graphite* (sheets of regular hexagons), *diamond* (three-dimensional tetrahedral network), *buckminsterfullerene* (spheres of 60 atoms), and a variety of *amorphous* (Greek *a* + *morphe* for not + form) forms, such as soot, lamp black, charcoal, coke, etc. While graphite is a good conductor of electricity and diamond is the hardest substance known, none of these forms are *malleable* (Latin *malleare* for hammer) as are the metals. Different element forms are called **allotropes** (Greek *allos* + *tropos* for other + way).

nuclear charge and increasing number of (mutually repulsive) electrons causes the atoms of all elements to be of comparable size, of the order of 100 pm (picometers), or 1 **Angstrom** ($1 \text{ \AA} \equiv 10^{-10} \text{ m}$) in radius. Nevertheless, subtle changes in size are reflected in noticeable trends in chemical and physical properties.

The basic trends in atomic size in the periodic table, measured first experimentally in the Nineteenth Century, and calculated theoretically in the twentieth, are that *size decreases across any given row and increases down any column in the periodic table.*⁷ The electronic configurations of the atoms help explain observed atom size trends in the periodic table. Fig. 8.1 shows that a new shell is occupied for each row of the periodic table. Quantum mechanics relates the radius of the electron cloud to the shell quantum number - the higher the n , the larger the radius (cf. Fig. 7.5). For this reason the size of atoms increases going down any column in the periodic table. In a given row, however, the outer shell of electrons remains relatively constant, suggesting constant size in a given row. Yet the atomic size decreases going across a row to the right in the periodic table. The reason for this involves a balance between two opposing effects. One is due to the nuclear charge (Z), which increases going across the table to the right, increasing the attraction for the electron clouds surrounding the nucleus (cf. Eq. (7.8)). The greater the attraction, the closer the cloud is drawn to the nucleus. An opposing effect derives from the larger numbers of electrons in the same shell resulting in greater electron repulsions, causing an expansion of the electron cloud. Of the two, the nuclear attraction dominates over the electron repulsion, resulting in a net decrease in size moving to the right along a row of the periodic table, while the opposite is true going down a column. See Fig. 8.2.

⁷ One might wonder how the size of an invisible atom could be determined. There were two methods available to Nineteenth-Century scientists, one for solids (metals) and another for gases (non-metals).

The density of an element relates its mass to volume. Given the atomic mass, a value for Avogadro's number and knowledge of the way the atoms are packed in the solid state, the atomic volume can be calculated. See the Crystal Atom Radius Algorithm in Section 16.1.

Gases are non-ideal to the extent that they have finite size and experience interactions. The "excluded volume" term in van der Waals' non-ideal gas equation (i.e. the b parameter in the $V - nb$ term) gives a direct measure of the volume occupied by the gas molecules (see Section 15.7).

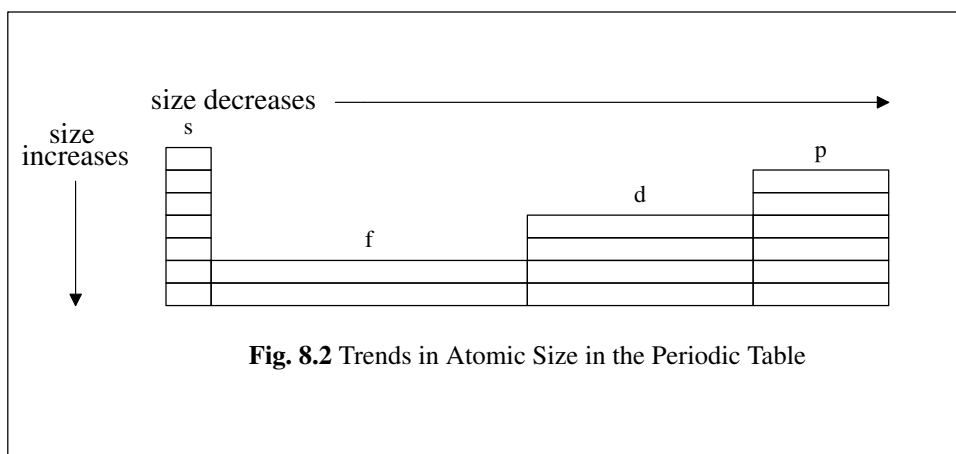


Fig. 8.2 Trends in Atomic Size in the Periodic Table

From the trends in atomic size, it is possible to explain other trends, ranging from atomic ionization energies to strengths of acids and bases. **Ionization** is the process of charging an atom or molecule, by removing or adding electrons. Since all neutral atoms have equal numbers of sub-atomic protons and electrons, the attraction between opposite charges prevents removal of electrons without an expenditure of energy. Hence, **ionization energies**, defined as the energy required to ionize an atom, are *always positive*, and always increase as more electrons are removed. The initial ionization process can be written symbolically as follows⁸

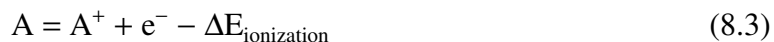


where $\Delta E_{\text{ionization}} > 0$ is the energy required to remove the electron e^- from the atom A (initial state), leaving ion A^+ (final state). Hence the notation $\Delta E_{\text{ionization}} \equiv E_{\text{products}} - E_{\text{reactants}} = E_{A^+ + e^-} - E_A$. Sometimes the notation I.E. (*ionization energy*) is used as an abbreviation for $\Delta E_{\text{ionization}}$. The arrow indicates the direction from initial state ("reactant") to final state ("product"). Since it is possible for an ion to become neutralized by capturing an electron, the "reverse reaction" exists as well:



⁸ Chemical reactions are expressed with a similar "equation" notation.

Equals signs are commonly used in place of arrows to represent reversible processes. The following equation is algebraically equivalent to Eq. (8.1) (with the arrow replaced by an equals sign), given the definition of Δ as the difference between initial and final states:



Ionization energies are dependent on how close the electrons are to the nuclei containing the protons of an atom. Generally, the smaller the atom, the larger the energy required to ionize it. This is consistent with Bohr's radius equation, Eq. (7.7) which shows that the energy of one-electron atoms decreases with increasing orbital quantum number n , where n represents the radius of the orbit (and size of a classical atom). Thus, ionization energy increases across any row of the periodic table (as size decreases) and decreases down any column (as size increases). Bohr's energy equation, Eq. (7.8), gives an estimate for the magnitudes of ionization energies: $E_n = \frac{-1312 Z^2}{n^2}$ kJ/mol. Since first ionization energies correspond to removing the outer electron to infinity, leaving a positive singly-charged ion, ionization energies are of the order of 1000 kJ/mol. This is only a crude estimate and requires solution to the wave equation for accurate values, which range from 376 kJ/mol for Cs to 2372 kJ/mol for He.

Attaching electrons to atoms and molecules is another story. Some elements resist adding extra electrons, such as the alkaline earths and noble gases with filled subshells of electrons. Electrons can be added to these elements, but a price in energy must be paid; the process is *endothermic* (energy goes into the system). For the alkaline earths, the larger the atom, the less endothermic the attachment process. Other elements, such as the halogens gain stability with additional electrons; for these the process is *exothermic* (energy goes out of the system). The energy *absorbed* in an electron attachment process is defined as the **electron affinity** of the atom (compare Eqs. (8.2) and (8.3)).



Unlike ionization energy, which is always a positive quantity, *electron affinities may be negative (in the case of electronegative nonmetals), positive (in the case of atoms with filled shell electronic configurations), or zero (in the case of nitrogen)*. Trends in electron affinity are not as regular as those in ionization energy, but roughly, the smaller the atom, the greater the electron affinity. First electron affinities range numerically between -348 kJ/mol for Cl (having positive affinity for electrons) and $+241$ kJ/mol for Be (having a negative affinity for electrons).

Note that since electrons are repulsive with each other, removing electrons from an atom decreases its size, while adding electrons to an atom increases its size. This explains general trends in successive ionization energies. For example, $\Delta E_{\text{ionization}}$ for Na, Na^+ and Na^{2+} equals 497, 4560 and 6900 KJ/mol, respectively, while $\Delta E_{\text{ionization}}$ for O and O^- are -142 and +780 KJ/mol, respectively.

Linus Pauling suggested that the average of the magnitudes of first ionization energy and electron affinity of an atom, which he coined **electronegativity**, is a measure of the attraction (greediness) for electrons of atoms bound in molecules. This quantity increases diagonally from cesium to fluorine, which is the most electronegative element. A number of mathematical formulas for electronegativity have been proposed, but the concept remains essentially qualitative. Electronegativity formulas are often scaled to range between about 1 for cesium to 4 for fluorine. In the second row of the periodic table, for instance, electronegativity increases by one-half unit per column from 1.0 for Li to 4.0 for fluorine, with boron at 2.0 representing the transition between metals and non-metals.

The discussion on trends in properties in the periodic table may be organized into a simple procedure:

The Periodic Trend Heuristic

Purpose: To determine trends in atomic properties.

Procedure:

1. Recognize that size decreases across rows and increases down columns of the periodic table (Fig. 8.2).
2. Use trends in size to predict trends in other properties. Smaller atoms have higher ionization energy, electronegativity and usually higher electron affinity unless the subshell is filled.

Example 8.7 Predict trends in density, melting point and bonding among the alkali metals.

1. The sizes of the alkali metals, Li, Na, K, Rb, Cs increase down the periodic table.

2. Density is the ratio of mass to volume. Both size and mass increase down a column of the periodic table. However, since the size of atoms does not increase as dramatically as the mass (recall all atoms are of roughly comparable size) density should increase down the alkali metal column. The facts are that while the atomic mass of cesium is 19 times larger than that of lithium (133 vs 7 amu), the radius of Cs is only about double that of Li (2.3 vs 1.2 Angstroms).

Melting is a loosening of the bonds which hold the atoms in position in the solid state. Bonds result from attraction of negative electrons on one atom to positive nuclei on others. The larger the atom, the weaker the attractions to its neighbors; hence less thermal energy (lower temperature) is needed to fluidize the atoms.

As ionic size increases, charge density decreases and attraction to neighboring atoms (bonding strength) is expected to decrease. This can be measured by determining the energy to separate metal ions from compounds, called *lattice energy*. Lattice energy should decrease with increasing size.

Trends in radius, density, melting point, ionization energy and metal-fluoride lattice energy all follow the predictions, as the table below shows.

Table 8.1 Experimental Properties of the Alkali Metals

<i>Element</i>	<i>Radius (Angstroms)</i>	<i>Density₃ (g/cm³)</i>	<i>Melting Point (°C)</i>	<i>Ionization Energy (kJ/mol)</i>	<i>Fluoride Lattice Energy (kJ/mol)</i>
Li	1.23	0.53	179	520	1036
Na	1.54	0.97	98	496	923
K	2.03	0.86	63	419	821
Rb	2.16	1.53	39	403	785
Cs	2.35	1.87	28	376	740

Summary

Chemical properties are to a first approximation dependent on the outer shell, or valence electrons. Elements with similar valence electronic configurations have similar chemical properties. Similar valence electronic configurations occur periodically with increasing atomic number. The periodic table is an arrangement of the elements by increasing atomic

number, with columns of families of elements with similar valence electronic configurations (and properties).

PERIODIC TABLE EXERCISES

1. Compare the average atomic masses of the nearest neighbors to germanium in the periodic table in the horizontal, vertical and diagonal directions.
2. Compare the average densities of the nearest neighbors to germanium in the periodic table in the horizontal and vertical directions. Densities: Si = 2.33 g/mL, Sn = 7.37 g/mL, Ga = 6.10 g/mL, As = 5.73 g/mL.
3. What do C, Si, Ge, Sn and Pb have in common regarding their electronic configurations?
4. What is the valence of arsenic?
5. Knowing that elemental carbon has two common forms, diamond and graphite, what can be said about silicon?
6. Hydrogen sulfide is poisonous. What can be inferred about hydrogen selenide?
7. How are the densities of sodium, magnesium and aluminum related?
8. Compare the trends in electron affinities of the alkali metals and halogens.

PERIODIC TABLE EXERCISE HINTS

1. An average is the sum of the values divided by the number of values.
2. Locate the neighbors to Sc in the periodic table.
3. Note where they are in the periodic table.
4. Arsenic is a representative element in family number five in the periodic table.
5. Silicon is in the same chemical family as carbon.
6. Selenium is in the same chemical family as sulfur.
7. How are sodium, magnesium and aluminum related in the periodic table?

8. Electron affinity should be related to size.