18. The Periodic Table

There are over 100 elements, each characterized by a different number of protons and a corresponding number of electrons in orbitals centered on the nucleus. Chemistry would be overwhelmingly complex if all elements were completely different from each other. But they are not. There are patterns in their behavior, some of which were recognized long before Rutherford, Bohr, and Schrödinger developed models for the atom.

Dmitri Ivanovitch Mendeleev (1834-1907) puzzled over patterns in the properties of the elements in 1871, the year of Rutherford’s birth. Mendeleev was born in Siberia. He studied science at St. Petersburg, where he graduated in chemistry in 1856. He became a professor of chemistry at St. Petersbur in 1863, and in 1870 he published his influential book, The Principles of Chemistry. By Mendeleev’s time it was well-known that some elements could be grouped into “families” that shared common characteristics such as ability to combine with other materials. The so-called alkali metals (lithium, sodium, potassium, rubidium, and cesium) are all metals that react with water—violently under some conditions. The halogens (fluorine, chlorine, bromine, and iodine) are all active salt-forming elements. In contrast, the noble gases (helium, neon, argon, and krypton) are inert and do not react spontaneously with other substances. In 1869 when Mendeleev began to publish his work, such families were well recognized, although a few of the member elements had not yet been discovered.

Mendeleev knew nothing of nuclei, protons, neutrons, or electrons—all yet to be discovered—but he did know the relative mass of the elements. He set out to see if there were some patterns in characteristics that could be connected to increasing atomic mass of the atoms. Today we use atomic number (the number of protons) rather than atomic masses to put elements in their proper order, but Mendeleev could not do so because it was still an unknown concept in his time.

What is meant by “atomic mass?” Each atom of an element has a mass that is primarily determined by the combined number of neutrons and protons in the nucleus (the rather light electrons make only a small contribution). The number of nucleons (protons plus neutrons) is the mass number. We often express the masses of atoms relative to the mass of carbon-12 (taken to weigh exactly 12 atomic mass units). When compared to carbon-12 the mass of the element is said to be expressed in atomic mass units.

What Mendeleev discovered was that by arranging the lighter elements in order of increasing atomic mass, the elements of families were always separated by about seven intervening elements. An analogy may help you to understand. Consider a heap of books, a mixture of art, science, philosophy, history, and so forth. Weigh each book and arrange the books in order of increasing weight. Take the eight lightest books and place them on the top shelf in order of increasing weight. Place the eight next lightest on the second shelf in like manner, and continue this process with all the books. Imagine your surprise if you then discover that the first book on each shelf is a science book, the second on each shelf an art book, the third on each shelf a philosophy book, and so on. You would have discovered a puzzling periodic law analogous to what Mendeleev discovered for the elements. In Mendeleev’s day he had no greater understanding why this should occur for elements than you would have for the books.

Periodic Patterns

You can see periodic patterns yourself if you look at some characteristics of the elements. Since the atomic masses were known, and since densities for solid elements were known, one could calculate atomic volumes:

$$\text{density} = \frac{\text{atomic weight}}{\text{atomic volume}}.$$  

Solving for atomic volume,

$$\text{atomic volume} = \frac{\text{atomic weight}}{\text{density}}.$$  

When the atomic volumes are plotted against atomic masses as in Figure 18.1, recurring peaks emerge. By counting the elements between the peaks at lithium (Li), at sodium (Na), and at potassium (K), you can see the periodic pattern with seven intervening atoms between
the peaks. Figure 18.2 shows modern data for the sizes of atoms, illustrating the trends in atomic diameters. In general, the atoms get larger moving down a column and they get smaller moving from left to right.

We can also see the periodic pattern in data that were unknown to Mendeleev. When an electron becomes a part of an atom, it falls into a kind of “energy well.” In other words, the electron loses energy by emitting a photon and sets up an orbital pattern around the nucleus. The electron now has less energy than it had when it was free. To become free again, the electron must regain its lost energy. We can, for example,
put the atom in an environment where its electrons can be bumped by other atoms or by free electrons. The energy required to remove the highest energy electron of an atom is called the “first ionization energy.” The first ionization energies for the light elements are plotted in Figure 18.3, but this time as a function of atomic number. The periodic pattern, eight elements wide, is quite striking. Ionization energies for the lighter elements are tabulated in Table 18.1. The second ionization energy is the energy required to free a second electron after the first one has already been removed.

Certain properties of elements correlate well with ionization energy. For example, all of the elements whose ionization energies are below 8 electron volts are metals (shiny, electrically conducting, and malleable). Each element whose ionization energy is above 10 electron volts lacks these characteristics and is called a nonmetal. Examples are carbon (charcoal); nitrogen and oxygen (the major gases in air); and helium, neon, and argon (transparent gases used in colored “neon” signs). Apparently there is a correlation between the periodic pattern and the appearance, natural state, and chemical aggressiveness of elements.

Mendeleev observed that even within families there were gradual changes as well. But the periodic pattern was still most striking, and he was so convinced of its validity that he made a bold and daring assertion to explain a gap in his table. “An element is wanting,” he said. But knowing the other members of its proposed family and knowing the trends of the Periodic Law, he could say in 1871:

The following are the properties which this element should have on the basis of the known properties of silicon, tin, zinc and arsenic. . . . the density of [the element] will be about 5.5 . . . it forms a higher oxide . . . and [the oxide] will have a density of about 4.7.

He went on to predict some of the compounds it would form: it would form volatile organometallic compounds that boil at about 160 °C; it would form a volatile liquid chloride which would boil at about 90°C and have a density of about 1.9; it would form compounds with sulfur that would be soluble in ammonium sulfide. And on and on he went. In 1887, C. Winckler of Freiburg discovered germanium, an element with almost exactly Mendeleev’s predicted properties.

Mendeleev went on to do a similar thing for the then-unknown elements gallium and scandium. Students thronged to his lectures, although many of them could not fully understand him.

Without knowing why the periodic patterns existed and with many of the elements yet undiscovered, Mendeleev and others were still able to arrange the elements into a table so that members of families appear as

![Figure 18.3. First ionization energies for several light elements. Note the low ionization energies (easy ionization) of the metals lithium, sodium, and potassium. The noble gases—helium, neon, and argon—have high ionization energies. (The ionization energies are expressed in the energy unit of the electron volt, which is the energy an electron acquires in accelerating from rest to a potential of 1 volt.)](image)

Table 18.1. Ionization energies in electron volts.

<table>
<thead>
<tr>
<th>ELEMENT</th>
<th>First</th>
<th>Second</th>
<th>Third</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>13.6</td>
<td></td>
<td></td>
</tr>
<tr>
<td>He</td>
<td>24.6</td>
<td>54.4</td>
<td></td>
</tr>
<tr>
<td>Li</td>
<td>5.4</td>
<td>75.6</td>
<td>122.5</td>
</tr>
<tr>
<td>Be</td>
<td>9.3</td>
<td>18.2</td>
<td>153.9</td>
</tr>
<tr>
<td>B</td>
<td>8.3</td>
<td>25.2</td>
<td>37.9</td>
</tr>
<tr>
<td>C</td>
<td>11.3</td>
<td>24.4</td>
<td>47.9</td>
</tr>
<tr>
<td>N</td>
<td>14.5</td>
<td>29.6</td>
<td>47.4</td>
</tr>
<tr>
<td>O</td>
<td>13.6</td>
<td>35.1</td>
<td>54.9</td>
</tr>
<tr>
<td>F</td>
<td>17.4</td>
<td>35.0</td>
<td>62.7</td>
</tr>
<tr>
<td>Ne</td>
<td>21.6</td>
<td>41.0</td>
<td>63.4</td>
</tr>
<tr>
<td>Na</td>
<td>5.1</td>
<td>47.3</td>
<td>71.6</td>
</tr>
<tr>
<td>Mg</td>
<td>7.6</td>
<td>15.0</td>
<td>80.1</td>
</tr>
</tbody>
</table>
the vertical columns of the table. The modern version of the Periodic Table can be seen in Appendix C. The elements in column IA are the family of alkali metals. The noble gases all appear in the far right column. The halogens are in the column labeled VIIA. Each small box contains the symbol of a different element (H for hydrogen, He for helium, etc.), its atomic number (in the upper left-hand corner), and the atomic mass (at bottom center). The elements are arranged according to atomic number rather than mass, but in only a few instances do they differ from Mendeleev’s ordering.

The grand accomplishment of science in the latter half of the 19th century was for Mendeleev and others to create the Periodic Table. In honor of Mendeleev’s accomplishment, element number 101, Md, is named mendelevium. It would be the grand accomplishment of Rutherford, Bohr, and others in the opening decades of the 20th century to explain why the mysterious periodic patterns exist.

How then does the Wave Model explain Mendeleev’s puzzling patterns? Figure 17.7 shows that the noble gases (helium, neon, argon) are just those atoms for which the uppermost occupied shell (or s and p subshells) is exactly filled. Helium with its two electrons just fills Shell 1. Neon adds 8 electrons (2 in s-orbitals, 6 in p-orbitals) to just fill Shell 2. Argon adds eight more to fill the s and p subshells of Shell 3. Each time a noble gas is encountered, which fills a shell or fills combined s and p subshells, the Periodic Table moves to a new row or “period” (see Fig. 18.4).

Figure 18.4. The shells of the atom. The element \( _2^3 \)He just fills Shell 1, the element \( _{10}^{18} \)Ne just fills Shell 2, and the element \( _{18}^{36} \)Ar just fills the combined s and p subshells of Shell 3.

Shell 3, however, should properly have electrons in the d-orbitals before it is completely filled. Nature’s pattern breaks at this point in building the atoms. Rather than completing Shell 3, nature begins to fill the fourth shell. The s-orbital of the fourth shell is lower in energy than the d-orbitals of the third shell, so potassium and calcium are formed before the pattern is resumed to fill the ten vacancies of the d-orbitals in the third shell. This break in the pattern emphasizes an important point: the patterns of the Periodic Table are striking, but they are not perfectly simple.

In the Periodic Table of Figure 18.2, the two columns on the left labeled IA and IIA (Li-Be, Na-Mg, K-Ca, etc.) contain elements in which the highest energy electrons are in s-orbitals. The block of six columns on the right (B-C-N-O-F-Ne; Al-Si-P-Si-Cl-Ar, etc.) are elements in which the highest energy electrons are in p-orbitals. The block of ten in roughly the center of the chart (Sc-Ti-V-Cr-Mn-Fe-Co-Ni-Cu-Zn, etc.) contains elements in which the next consecutive electrons are in d-orbitals.

Another periodic pattern of the elements is their chemical behavior, which for most atoms depends only on those electrons in the outermost shell. The electrons in the outermost shell are called valence electrons. Elements in the same column of the Periodic Table have the same number of valence electrons and belong to a family with common chemical properties. All the elements in column IA are elements that have one electron more than the number required to fill a shell or combined s and p subshells. These elements are said to have one valence electron. The elements in column IIA have two valence electrons. The elements in column IIIA have three valence electrons, and so on.

Mendeleev’s patterns are simply the result of the filling of the shells. The existence of the shells and their numbers of available energy states are dictated by Schrödinger’s wave equation.

Historical Perspectives

Thales of Mile tus (ca. 624-546 B.C.) thought that water was the elemental substance from which all else was made. Others believed it to be air and fire. Empedocles the Sicilian (ca. 492-425 B.C.) compromised with a pluralistic basis of elements: earth, water, air, and fire. Aristotle added quintessence to account for the stuff of the celestial objects. Some of these ideas were carried to Egypt with the conquering army of Alexander the Great where they were mixed with the practical Egyptian art of metallurgy. From there the Muslims nursed the ideas into the 7th century and eventually carried them back to Europe in the 13th century as “alchemy.”

Alchemy was a curious mix of science, religion, and mysticism. Legend traced its origins alternately to the Egyptian god Hermes; to the fallen angels spoken of in the Book of Enoch who revealed to men the knowledge of gold and silver and the power of herbs; or to Moses and Aaron, who were to have received it directly from God.
Inorganic material was thought of as living beings, with a body of matter and a spirit that determined its characteristics and properties. Metals could be ennobled by undergoing a process of death and resurrection. The spirit could be separated from the matter by vaporization (heating) and some vapors could be liquefied by cooling them (distillation). For example, alcohol was the spirit of wine and beer and is often referred to as “spirits” to the present day. Liquids obtained by distillation (the spirit) were highly potent and active agents containing the concentrated essence of a material. Thus, the spirit of noble metals could be transferred to other materials (such as base metals) and give them new life and the properties of the new spirit within them. Since mercury was the only metal that could be distilled, it was considered the progenitor of the metals and the origin of all things. For example, mercury could be rubbed on base metal to give it the appearance of silver; therefore, mercury contained something of the spirit of silver. Metals were generally taken to be a product of the female principle of mercury and the male principle of sulfur.

Paracelsus (1493-1541) continued the alchemist tradition, but added salt to sulfur and mercury to complete a triad of elemental substances. He added salt out of the religious conviction that the world followed the Trinity in being organized in triads. Paracelsus, however, applied his theories to human health by treating illness as an imbalance in the three principles within the human body. He regarded his system as a revelation that would restore the purity of medicine in much the same way the religious reformers of his time regarded their theologies as revelations to restore primitive Christianity. Paracelsus, with the alchemists, largely rejected reason and sensory perception as ways of knowing and therefore made virtually no scientific progress.

Robert Boyle (1627-1691), an alchemist, began to break with tradition and to usher in a science of chemistry in place of alchemy. He published his ideas in his book, The Skeptical Chemist, in 1661. Boyle rejected the earlier metaphysical notions of elements and replaced them with an operational definition: An element is any substance that cannot be separated into different components by any known methods. By this definition, over the next century some of the old substances (gold, silver, mercury, copper) were recognized as elements. To these were added hydrogen (Henry Cavendish, ca. 1750), oxygen (Joseph Priestley, 1774), nitrogen (Daniel Rutherford, K. W. Scheele, ca. 1772), and others in a fairly constant stream, until the last of the 88 natural elements (rhenium) was identified in 1925. The highest atomic number among the naturally occurring elements is 92 (uranium), but four elements (43, 61, 85, 87) are so rare as to be considered “missing.” Hence, we speak of 88 natural elements.

Following the definition of an element, Antoine-Laurent Lavoisier (1743-1794) produced the next important insight. Lavoisier was a brilliant aristocrat who was beheaded in the French Revolution for having “added water to the people’s tobacco.” Despite his troubles with the Revolution, he was a multifaceted genius who demonstrated the advantages of scientific agriculture and planned the improvement of social and economic conditions by means of savings banks, insurance societies, canals, and so forth. Still, he became involved with some activities that made him an object of suspicion. Joseph-Louis Lagrange observed following his death: “It required only a moment to sever that head, and perhaps a century will not be sufficient to produce another like it.”

Before his execution Lavoisier came to grips with a puzzle: Why was it that some substances (wood, for example) lost weight when burned, while others (phosphorus) actually gained weight? He studied combustion with carefully measured quantities of gases in closed containers and discovered what came to be known as the Law of Conservation of Mass: The total quantity of matter within the system remains constant (ca. 1789). Wood appears to lose weight when burned because the gases from the combustion (carbon monoxide, carbon dioxide, and steam) escape and are not accounted for. Phosphorus appears to gain weight because oxygen from the surrounding atmosphere combines with phosphorus to form a solid oxide. When all gases are properly accounted for, no change in mass is detected.

Next came Joseph-Louis Proust (1755-1826) and Proust’s Law: Elements always combine to form compounds in certain definite proportions by weight (ca. 1800). For example, in forming water exactly 8 grams of oxygen combine with each gram of hydrogen.

Proust’s Law allowed John Dalton (1766-1844) to associate the idea of an atom with the concept of an element: An element is a substance composed of only one kind of atom: atoms of a given substance are identical (ca. 1803). Dalton also assumed that when two or more atoms combine to form a molecule, they do so with the smallest possible number of atoms. Thus, he assumed one hydrogen atom combines with one oxygen atom to form water. Using Proust’s Law, Dalton was forced to conclude that oxygen atoms weigh eight times as much as hydrogen atoms. He drew up a table of relative atomic masses for the known elements that was largely wrong, but the idea of an atom was firmly established.

While Dalton was finding great significance in the ratios of combining weights, Frenchman Joseph-Louis Gay-Lussac (1778-1850) was concentrating on combining volumes of gases (taken at the same temperature and pressure). Dalton, supposing water to form from one atom of hydrogen and one atom of oxygen, predicted that the two atoms would bind closely together. When a volume of each element was combined, he thought one volume of water vapor would result (see Fig. 18.5). But in 1808 Gay-Lussac found that it wasn’t so at all.
Instead, 2 volumes of hydrogen gas added to 1 volume of oxygen gas resulted in 2 volumes of water vapor. Dalton was flabbergasted. He rejected Gay-Lussac’s results with some vehemence. He accused the Frenchman of carelessness and performed his own crude experiments and found they disagreed.

The missing piece to the puzzle was supplied by Lorenzo Romano Amedeo Carlo Avogadro di Quaregna e di Ceretto (1776-1856). Amedeo Avogadro made a conceptual leap that was ahead of its time. He said that perhaps even the “atoms” of gases are not necessarily single atoms, but rather groups of atoms stuck together (1811). Such a group is called a molecule. He speculated that oxygen is made of two atoms stuck together. The problem with this assumption—and one reason that others hadn’t proposed it earlier—was that if two atoms of an element coalesced, why not three or four or five? Why, in fact, would not all of the atoms of the container of gas stick together into one horrendous molecule? Avogadro had no answer, but made the leap anyway. His second conjecture is known as Avogadro’s Hypothesis: Equal volumes of all gases (at the same temperature and pressure), whether elements, compounds, or mixtures, contain equal numbers of molecules (1811) (see Fig. 18.6.). (Each molecule in a compound contains atoms of at least two different elements. Any two or more substances can be loosely combined into a mixture that can be mechanically “unmixed”.)

Avogadro’s brilliant breakthrough was largely ignored for another 50 years until Stanislao Cannizzaro revived it in 1860 at the First International Chemical Congress at Karlsruhe, Germany. Professor Lothar Meyer later remarked of Cannizzaro’s presentation: “It was as though scales fell from my eyes, doubt vanished, and was replaced by a feeling of peaceful certainty.” Chemists, too, have their moments of revelation.

**STUDY GUIDE**

Chapter 18: The Periodic Table

A. FUNDAMENTAL PRINCIPLES

1. Law of Conservation of Mass (as expressed by Lavoisier [ca. 1789] for chemical reactions): The total quantity of matter within the system remains constant. See also Chapters 7 and 9.

2. The Wave-Particle Duality of Matter: See Chapter 16.


B. MODELS, IDEAS, QUESTIONS, OR APPLICATIONS

1. When studying the elements, what is observed about their volumes, ionization energies, and chemical properties when they are organized according to atomic number?

2. What happened in the early part of this century that brought greater unity to science?

3. How does the Wave Model of the atom explain the “periodicity” that led to the Periodic Table?

C. GLOSSARY

1. Avogadro’s Hypothesis: Equal volumes of all gases (at the same temperature and pressure), whether elements, compounds, or mixtures, contain equal numbers of molecules.

2. Atomic Mass: Although sometimes referred to as Atomic Weight, what is meant is the mass per unit atom of a sample of substance. Ordinarily the atomic masses are expressed in units in which the mass of carbon is exactly 12.0.

3. Atomic Number: See Chapter 17.
4. **Atomic Volume**: The volume of an atom.
5. **Element**: A substance made up of atoms all having the same number of protons.
6. **Families (Groups)**: Families (or Groups) are elements in the Periodic Table which are placed in the same vertical column. Elements in a given family have similar chemical behavior and properties.
7. **Ionization Energy**: The amount of energy which an electron must acquire to be removed from an atom.
8. **Mass Number**: See Chapter 17.
9. **Metals**: The class of elements occurring to the left of the Periodic Table and usually characterized as shiny or whitish, malleable, solid, good conductors of electricity and heat, and chemically active.
10. **Nonmetals**: The class of elements occurring to the upper right of the Periodic Table and lacking at least one of the characteristics of metals.
11. **Periods**: Elements that occur in the same row of the Periodic Table are said to belong to the same period. Elements in a period exhibit systematically changing chemical behavior which is best seen in the formulas of the compounds the elements in the period form with a given other element.
12. **Proust’s Law**: Elements always combine to form compounds in certain definite proportions by weight.
13. **Valence Electrons**: The name given to the electrons occupying the highest unfilled shell in an atom.

**D. FOCUS QUESTIONS**

1. Describe three important patterns of the properties of the chemical elements that correlate to their placement on the Periodic Table. What are the main elements of the Wave Model of the Atom (including a statement of two fundamental principles upon which it is built)? How does the Wave Model of the Atom explain the patterns?

**E. EXERCISES**

18.1. Which would you expect to be the larger atom, \( _{37}^{79}\text{Rb} \) or \( _{17}^{35}\text{Cl} \)?

18.2. Which two of the following atoms are most nearly the same size: \( _{1}^{3}\text{F}, _{30}^{63}\text{Zn}, _{47}^{101}\text{Ag}, \) or \( _{56}^{112}\text{Ba} \)?

18.3. Which would be more dense, \( _{13}^{24}\text{Mg} \) (magnesium used in the original mag wheels) or \( _{79}^{195}\text{Au} \) (gold)?

18.4. Which is the smaller neutral atom, \( _{38}^{87}\text{Sr} \) or \( _{16}^{32}\text{S} \)? (You should be able to answer this question on the basis of position in the Periodic Table.)

18.5. Which is more dense, \( _{22}^{44}\text{Ti} \) or \( _{79}^{195}\text{Au} \)?

18.6. Observe the distance from the highest energy electron to the top of the well (0 electron volts) in Figure 17.6. This would be a measure of ionization energy. Which of the four atoms shown in Figure 17.6 has the lowest ionization energy (shortest distance to the top of the well)? Check your answers in Table 18.1.

18.7. Using Table 18.1 and the rule that metals have ionization energies of less than 8 electron volts and nonmetals of more than 10 electron volts, mark each of the following as metal, nonmetal, or cannot tell:

(a) \( _{1}^{1}\text{H} \), hydrogen
(b) \( _{4}^{9}\text{Li} \), lithium
(c) \( _{6}^{15}\text{B} \), boron
(d) \( _{10}^{18}\text{Ne} \), neon
(e) \( _{12}^{24}\text{Mg} \), magnesium

18.8. From Figure 17.6 state which electron Li is likely to lose to become \( \text{Li}^{+} \). Do the same for the two electrons Be loses to become \( \text{Be}^{2+} \).

18.9. Why is the second ionization energy of Li so much higher than the first? (See Table 18.1 for ionization energies, and see Figure 17.6 for a hint about the explanation.)

18.10. Which of the following pairs of elements will be most like \( _{13}^{33}\text{Al} \) and \( _{16}^{32}\text{S} \)? (a) \( _{46}^{117}\text{Pd} \) and \( _{82}^{209}\text{Pb} \), (b) \( _{14}^{30}\text{Si} \) and \( _{16}^{32}\text{S} \), (c) \( _{17}^{35}\text{Cl} \) and \( _{35}^{79}\text{Br} \).

18.11. Which element is most similar to \( _{18}^{38}\text{Ar} \)?

18.12. Technetium, \( _{43}^{99}\text{Tc} \), does not exist naturally on the earth and is prepared by nuclear bombardment. If you were looking for it, would you expect to find it as a solid, liquid, gas, metal, or nonmetal? (This element inhibits the corrosion of steel remarkably well. At one time it cost $2,800 per gram, but is now below $100 per gram.)

18.13. Which gaseous element has the largest number of protons in each atom?

18.14. On the basis of its neighbors, describe the characteristics of \( _{46}^{109}\text{Pd} \).

18.15. Which two elements are most like \( _{35}^{80}\text{Br} \)?

18.16. Compared to chromium, \( _{24}^{52}\text{Cr} \), what can you say about molybdenum, \( _{42}^{94}\text{Mo} \)? Compare density and metallic nature.
18.17. Draw lines between atoms which belong to the same family:

\[ \begin{array}{ll}
20^{\text{Ca}} & 18^{\text{Ar}} \\
3^{\text{Li}} & 12^{\text{Mg}} \\
35^{\text{Br}} & 33^{\text{As}} \\
10^{\text{Ne}} & 11^{\text{Na}} \\
7^{\text{N}} & 9^{\text{F}}
\end{array} \]

18.18. Which of the following are nonmetals?
\[ 16^{\text{S}}, \ 25^{\text{Cs}}, \ 48^{\text{Cd}}, \ 15^{\text{P}}, \ 23^{\text{V}} \]

18.19. Which gaseous element has the smallest number of electrons per atom?

18.20. List Fe, Rh, and Pt in order of increasing density. (They all have about the same diameter.)
(a) Pt Rh Fe
(b) Pt Fe Rh
(c) Rh Fe Pt
(d) Rh Pt Fe
(e) Fe Rh Pt

18.21. List the following atoms in order of increasing diameter: nitrogen, oxygen, phosphorus:
(a) N O P
(b) N P O
(c) P N O
(d) P O N
(e) O N P

18.22. The atomic mass of a potassium atom is about 39. The atom has
(a) 39 protons
(b) 39 neutrons
(c) 39 electrons
(d) 19 neutrons
(e) 19 protons

18.23. How many valence electrons (and in what orbitals) does oxygen have?
(a) 8 in p orbitals
(b) 6 in p orbitals
(c) 4 in p orbitals
(d) 4 in s orbitals
(e) 2 in s orbitals

18.24. Which atom has the lowest ionization energy?
(a) hydrogen
(b) helium
(c) lithium
(d) beryllium
(e) aluminum

18.25. Which of the following statements is false?
(a) Elements in a row progress from metal to nonmetal.
(b) Elements in a column are chemically similar.
(c) Li, Na, and K are chemically similar.
(d) Ne, Ar, and Kr are chemically active solids.
(e) F, Cl, and Br are chemically active gasses.

18.26. Which of the following are most chemically similar to Cl?
(a) F and Br
(b) Cu and Au
(c) Na and K
(d) P and S
(e) Br and Kr