6 Relationships and Patterns in Chemistry

This chapter focuses on the following AP Big Idea from the College Board:

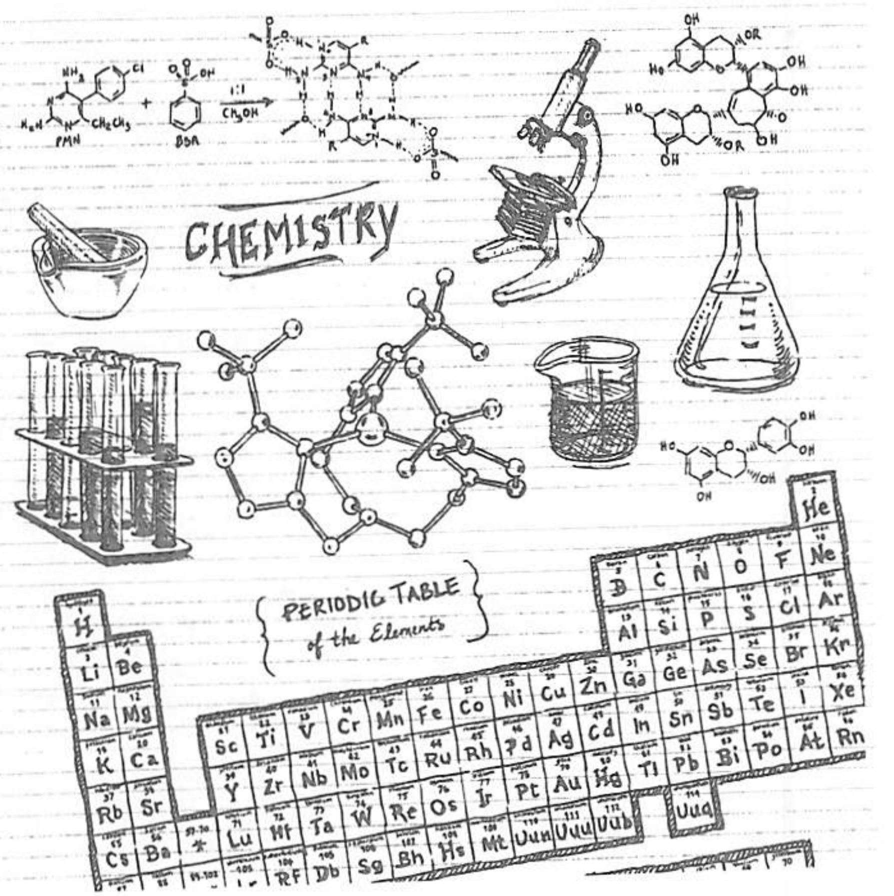
Big Idea 1: The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identify in chemical reactions.
Big Idea 2: Chemical and physical properties of materials can be explained by the structure and the arrangement of atoms, ions, or molecules and the forces between them.

By the end of this chapter, you should be able to do the following:

- Describe the development of the modern periodic table
- Draw conclusions about the similarities and trends in the properties of elements, with reference to the periodic table
- Justify chemical and physical properties in terms of electron population
- Demonstrate knowledge of various types of chemical bonding
- Apply understanding of bonding to create formulae and Lewis structures

By the end of this chapter you should know the meaning of the following **key terms**:

- alkali metals
- alkaline earth metals
- atomic radius
- covalent bonding
- electrical conductivity
- electron dot diagram
- halogens
- ionic bonding
- ionization energy
- Lewis structure
- melting point
- metal
- metalloid
- mole
- noble gases
- non-metal
- polarity
- transition metals
- valence electrons



6.1 The Development of the Periodic Table

Warm Up

1. Can you suggest the meaning of the word "periodic" in the term "periodic table"?

2. On what basis are the elements arranged in the modern periodic table?

3. What is true about chemical elements that appear in the same vertical column in the table?

Discovering an Elemental Order

Science in general and chemistry in particular exist because our human species has always had an insatiable desire to make sense of the world around us. We have relentlessly sought to explain nature's phenomena, to solve her mysteries, and to discover her order and logic by deciphering the events and objects we encounter.

One of the most important and successful examples of such efforts is the development of the periodic table of the elements. This single document is arguably more valuable to chemists, and perhaps even society itself, than any piece of equipment, wonder drug, or process ever invented.

The periodic table had its beginnings in the early part of the 1800s. By 1817, 52 elements had been discovered. Although some had been known since ancient times, many new elements were being discovered using the energy available from the electric battery invented by Volta in 1800. Researchers saw the need to organize those elements and the enormous amount of information gathered about them into some kind of meaningful form. German scientist Johann Dobereiner noticed similarities within several groups of three elements such as chlorine, bromine, and iodine, which he called "triads." These similarities gave chemists some evidence that an organizational scheme was at least possible.

In 1857, English chemist William Odling proposed that the elements could be divided into groups based on their chemical and physical properties. Many of those groups actually resemble the vertical columns in the periodic table today.

In 1862, French geologist Alexandre-Emile de Chancourtois arranged the elements by increasing atomic mass. He noted that elements with similar properties seemed to occur at regular intervals. He devised a spiral graph with the elements arranged onto a cylinder such that similar elements lined up vertically. When his paper was published, however, it was largely ignored by chemists. Unfortunately, de Chancourtois had left out the graph, which made the paper hard to understand. As well, he had written the paper from a geologist's rather than from a chemist's perspective.

In 1864, English chemist John Newlands noticed that similar properties seemed to repeat every eighth element in much the same way that the notes of a musical scale repeat every eighth tone. Newlands called this the "law of octaves" and it resulted in several similar elements being grouped together, but with limited success.

The Periodic Law

The most successful organization of the elements was arrived at independently and almost at the same time by the German chemist Julius Lothar Meyer and the Russian chemist Dmitri Ivanovich Mendeleev. The chemical community, however, ultimately awarded Mendeleev the majority of the credit.

In a paper submitted to the Russian Chemical Society in March 1869, Mendeleev arranged the elements into a **periodic table** and proposed the **periodic law**, which can be stated as follows:

If elements are arranged in order of increasing atomic mass, a pattern can be seen in which similar properties recur on a regular or *periodic* basis.

Notice that this is called a "law" rather than a "theory." Theories attempt to explain *why* relationships or phenomena exist, whereas laws simply identify *that* they exist. We will see if we can supply the "why" by the end of this section.

Dmitri Mendeleev received most of the credit for several reasons. First, Mendeleev's work was published a year before Meyer's. Second, Mendeleev chose to concentrate on the chemical properties of the elements, while Meyer focused mainly on physical properties. Third, and most importantly, Mendeleev chose to leave several blank spaces in his table where he predicted as-yet-undiscovered elements with specific properties would eventually be placed. Because he had arranged similar elements vertically in his table, the location of those blank spaces effectively identified the properties that those elements would have upon being discovered. When the elements were eventually discovered, the predictions turned out to be extremely accurate.

A version of Mendeleev's table, published in 1872 and showing 65 elements is shown in Figure 6.1.1.

Reiben	Gruppo I. — R^2O	Gruppo II. — RO	Gruppo III. — R^2O^3	Gruppo IV. RH^4 RO^2	Gruppo V. RH^3 R^2O^3	Gruppo VI. RH^3 RO^3	Gruppo VII. RH R^2O^7	Gruppo VIII. — RO^4
1	II = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,8	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	— = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fo = 56, Co=59, Ni = 59, Cu=63.
5	(Cu = 63)	Zn = 65	— = 68	— = 72	As = 75	So = 78	Br = 80	
6	Rb = 86	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	— = 100	Ru = 104, Rh=104, Pd = 106, Ag=108.
7	(Au = 199)	Cd = 112	In = 113	Sn = 118	Sb = 122	To = 125	J = 127	— — — —
8	Cs = 133	Bs = 137	?Di = 138	?Ce = 140	—	—	—	
9	(—)	—	—	—	—	—	—	
10	—	—	?Er = 178	?La = 180	Ta = 182	W = 184	—	Os = 195, Ir=197, Pt = 198, Au=199.
11	(Ag = 108)	Hg = 200	Tl = 204	Pb = 207	Bi = 208	—	—	— — — —
12	—	—	—	Th = 231	—	U = 240	—	

Figure 6.1.1 *The blank spaces marked with lines for elements with atomic masses 44, 68, 72, and 100 represent Mendeleev's belief that those elements would eventually be discovered and fit in the spaces. The symbols at the top of the columns (e.g., R^2O and RH^2) are molecular formulas written in the style of the 1800s. The letter "R" represents any element in the family and the formulas represent the probable hydrogen and oxygen compounds.*

An example of the accuracy of one of Mendeleev's predictions (and thus support for his periodic table) can be seen in Table 6.1.1. The table displays the properties predicted in 1871 for the element with atomic mass 72, compared to the actual properties for that element, called germanium, which was discovered in 1886.

Table 6.1.1 *Predicted and Observed Properties of Germanium*

Properties of Germanium	Predicted Properties in 1871	Observed Properties in 1886
Atomic mass	72	72.3
Density	5.5 g/cm^3	5.47 g/cm^3
Melting point	very high	960°C
Specific heat	0.31 J/g °C	0.31 J/g °C
Oxide formula	RO_2	GeO_2
Oxide density	4.7 g/cm^3	4.70 g/cm^3
Chloride formula	RCl_4	$GeCl_4$

Using the oxide formulas that he had proposed as a guide, Mendeleev also corrected the atomic masses of the elements beryllium, indium, and uranium.

Because of its obvious usefulness, Mendeleev's periodic table gained widespread acceptance among chemists. We should remember that the periodic table was constructed prior to any discoveries about the inner structure of the atom. The similarities of various elements were eventually explained based on the quantum mechanical description of their electron arrangements, but the table identified those regularities more than 50 years before they were understood!

Quick Check

1. Who was the first person to arrange the chemical elements into groups?

   ____

2. How did Newland's analogy to music apply to elemental properties?

   ____

3. How did blank spaces in Mendeleev's periodic table help it eventually gain acceptance?

   ____

   ____

The Modern Periodic Table — By the Numbers

Since its creation, the periodic table has undergone several changes. Many new elements have been discovered or synthesized since 1872, but the most significant modification occurred in 1913. Data gathered by the young British chemist Henry Moseley, combined with the discovery of isotopes, resulted in the elements of the periodic table being re-ordered according to their *atomic numbers* rather than their atomic masses.

If elements are arranged in order of increasing atomic number, a pattern can be seen in which similar properties recur on a regular or periodic basis.

The Modern Periodic Table

With the inclusion of element 117 in April 2010, the periodic table includes 118 elements, 92 of which occur naturally. Each element is assigned its own box in the table containing that element's one- or two-letter symbol, atomic number and sometimes, but not always, atomic mass. The boxes are arranged in order of increasing atomic number beginning at the top left with hydrogen and proceeding horizontally left-to-right. Because elements 113 to 118 have not yet been assigned permanent names, their temporary names simply reflect their atomic numbers. Figure 6.1.2 shows a simplified periodic table.

The horizontal rows in the periodic table are called **periods** and these are numbered beginning at the top from 1 down to 7. Each vertical column is called a **family** or **group** and these are labeled using one of two schemes. Many chemists prefer an older system, which assigns each column a number from 1 to 8 combined with either the letter A or B.

A newer system adopted by the international governing body of chemistry called IUPAC (International Union of Pure and Applied Chemistry) simply numbers each group from 1 through 18 and uses no letters. For ease of reference, we will use the IUPAC numbering system.

Because each element is a member of both a horizontal period and a vertical group, its position on the table can be specified in much the same way as a geographic location on a map can be identified using lines of longitude and latitude.

The two most important things to know about the periodic table are:

1. Elements in the same chemical family have similar chemical properties because they have similar outer electron configurations. (This is among the most important of all chemical concepts.)
2. As we move down a chemical family or across a period, there are regular changes in these properties called periodic trends. (These will be discussed in detail in the next section.)

Groups 1 and 2 on the left, and 13 through 18 on the right contain the **representative** or **main group elements**. Groups 3 through 12 are the **transition elements**. Within the transition elements are two horizontal series collectively called the **inner transition elements**. The first series, known as the **lanthanides**, fits between elements 56 and 72 in the 6th period and the second series, called the **actinides**, fits between elements 88 and 104 in the 7th period.

Representative Elements — Transition Elements — Representative Elements

Group → / Period ↓	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Inner Transition Elements

Figure 6.1.2 *This periodic table shows only the element's symbols and their atomic numbers.*

Consider the periodic table in Figure 6.1.2. The first distinction among the elements relating to their properties that we will discuss is their classification as metals, non-metals, or metalloids. The "staircase" line descending from group 13 down to group 16 separates metals on the left side from non-metals on the right.

About three-quarters of the elements are metals, including some main group elements and all of the transition and inner transition elements. Properties of **metals** include:

- solids at room temperature, except for mercury, which is a liquid
- generally shiny or lustrous when freshly cut or polished
- good conductors of heat and electricity
- generally malleable, which means they can be rolled or hammered into thin sheets
- generally ductile, which means they can be rolled or stretched into wires
- generally flexible as thin sheets or wires
- during chemical changes, tending to give up electrons relatively easily to form cations

The **non-metals** are located in the upper right portion of the table. Properties of non-metals include:

- usually gases or brittle solids at room temperature, except for liquid bromine
- solid non-metals ranging in appearance from dull or lustrous and translucent to opaque
- poor conductors of heat and electricity
- during chemical changes, tending to gain electrons from metals to form anions or share electrons with other non-metals.

Along the staircase line are several **metalloids** (also called **semi-metals**), including the elements boron, silicon, germanium, arsenic, antimony, tellurium, and polonium. These solid elements are semiconductors, which mean they are poor electrical conductors at room temperature, but become better conductors at higher temperatures. This is opposite to how metal conductivity varies with temperature. Metals become less conductive as temperature increases.

As we move right-to-left across a period and down a chemical family, the elements become more metallic. These changes involve both physical and chemical properties, and those and other trends will be discussed in the next section.

Quick Check

1. Fill in the missing spaces in the table below:

Element Name	Element Symbol	Group Number	Period Number	Metal/Non-metal/ Metalloid
	Si			
osmium				
		6	4	
	Mt			
antimony				metalloid
		17	5	

2. (a) Rearrange the following alphabetical list in order of least metallic to most metallic.
aluminum, cesium, chromium, fluorine, gallium, oxygen, sulphur, zirconium

____

(b) Based on your answer to (a) above, which of the eight elements would you expect to have the:

- greatest tendency to gain an electron? ____________________
- greatest tendency to lose an electron? ____________________

A Closer Look at the Periodic Table

It is the number and type of outermost electrons that are primarily responsible for an atom's chemistry. The members of a chemical family have similar numbers and types of outermost electrons. The outermost electrons that participate in chemical bonding are known as **valence electrons**. We will be discussing chemical bonding in more detail in a later section. For now, this section will introduce some families in the periodic table and the corresponding electron arrangements that will influence the bonding behaviour of their elements.

Group 1 — The Alkali Metals

Alkali metals are located on the far left side of the periodic table. This group includes lithium, sodium, potassium, rubidium, cesium, and the rare and radioactive francium. Alkali metals are all soft, silvery solids and the most reactive of all metals. The name of the group is based on the fact that the oxide compounds of the alkali metals dissolve in water to produce strongly basic solutions. They all corrode rapidly in air to a dull gray appearance, react vigorously with water to produce hydrogen gas, and readily form compounds with non-metals.

Alkali metal atoms have one valence electron in the s sublevel with the general electron configuration [noble gas]ns^1 where "*n*" represents the outer energy level. In their chemical reactions, they readily lose that outer electron to form 1+ cations and so assume the electron configuration of the previous noble gas. Notice in Table 6.1.2 that each of the alkali metals has the same number and type of valence electrons. Note also, that as we move down the group and the value of "*n*" increases, the outermost electron spends most of the time farther and farther from the nucleus so the atoms become larger.

Table 6.1.2 *Alkali Metal Electron Configurations*

Alkali Metal	Core Notation Configuration
Li	[He] $2s^1$
Na	[Ne] $3s^1$
K	[Ar] $4s^1$
Rb	[Kr] $5s^1$
Cs	[Xe] $6s^1$
Fr	[Rn] $7s^1$

Group 2 — The Alkaline Earth Metals

The **alkaline earth metals** are also silver-coloured reactive metals. Although they are not as reactive as the alkali metals, they readily form compounds with non-metals. Their oxides are also alkaline but unlike alkali compounds, some group 2 compounds have a low solubility in water. They have two valence electrons and have the general electron configuration [noble gas]ns^2. They will readily form 2+ cations by losing those two valence electrons and so will achieve the identical electron configuration of the nearest noble gas. Chemists call this becoming **isoelectronic** with the noble gas.

Group 17 — The Halogens

The family of **halogens** contains the most reactive elements and is the only one in which all three states of matter are represented. At room temperature, fluorine and chlorine are gases, bromine is a liquid, and iodine and the very rare and radioactive astatine are solids.

The elemental halogens exist as diatomic molecules and readily form compounds with metals, and also hydrogen, carbon, and other non-metals. The halogens possess seven valence electrons with two electrons in the outer s sublevel and five electrons in the p sublevel. They therefore have the general ns^2np^5 outer electron configuration. In their reactions with metals, halogens will typically gain a single electron forming 1– anions and acquiring a noble gas electron configuration. When they react with non-metals, they will often share valence electrons.

Group 18 — The Noble Gases

As the name suggests, the family of **noble gases** is generally unreactive, although compounds of argon, krypton, xenon, and radon have been prepared.

All of the noble gases, except helium, have filled s and p sublevels. As the s and p sublevels are an atom's outermost orbitals, atoms of neon down to radon have eight electrons in their outer charge clouds with the ns^2np^6 configuration. They are said to possess "stable octets." This particularly stable electron configuration explains the low reactivity of the noble gases. It also explains the electron configuration achieved by many cations and anions. For example, consider Table 6.1.3, showing elements from four different families. Note that the stable ion of each element has the same electron configuration as the noble gas nearest to it in the periodic table.

Table 6.1.3 *Electron Configurations of Noble Gases and Nearby Elements*

Element	Electron Configuration for Atom	Nearest Noble Gas	Symbol and Electron Configuration for Stable Ion	
oxygen	$[He]2s^22p^4$	neon	O^{2-}	$[He]2s^22p^6$
calcium	$[Ar]4s^2$	argon	Ca^{2+}	$[Ne]3s^23p^6$
arsenic	$[Ar]4s^23d^{10}4p^3$	krypton	As^{3-}	$[Ar]4s^23d^{10}4p^6$
cesium	$[Xe]6s^1$	xenon	Cs^+	$[Kr]5s^24d^{10}5p^6$

Transition Elements

The **transition elements** include groups 3 through 12 in the periodic table. All are metals and most are hard solids with high melting and boiling points. An explanation of their chemical behaviour is beyond the scope of this course, but it differs from the representative elements. For example, the transition elements show similarities within a given period as well as within a group. In addition, many transition elements form cations with multiple charges and those cations often form complex ions. A number of transition metal compounds have distinct and recognizable colours. The differences occur mainly because the last electrons added for transition metals are placed in inner d orbitals. Electrons in these orbitals are usually closer to the nucleus than in the outer s or p orbitals filled for the representative elements.

Quick Check

1. Identify the family number and name to which each of the following properties best apply:

Property	Family Number	Family Name
(a) reactive non-metals possessing seven valence electrons		
(b) reactive solids that form 2+ cations during reactions		
(c) invisible gases that are almost totally unreactive		
(d) soft, very reactive silvery solids with one valence electron		

2. Elements in the same family demonstrate similar chemical behaviour. Consider the following chemical formulas: $LiBr$, K_2O, Sr_3N_2, AlF_3, CaO, H_2S. Write chemical formulas for:

   (a) sodium iodide ______________ (d) rubidium sulphide ______________

   (b) barium sulphide ______________ (e) magnesium phosphide ______________

   (c) gallium chloride ______________ (f) hydrogen selenide ______________

3. Which two families contain the most reactive elements? Can you suggest a possible reason for this given their location on the periodic table?

   ____

   ____

A Quantum Mechanical View of the Periodic Table

Let us look a little more closely at the organization of the periodic table from a quantum mechanics perspective. First we'll look at three different aspects of the elements in the periodic table.

(1) Consider Figure 6.1.3 showing the chemical families separated into the four main blocks of the table. Notice that each block corresponds to one of the four different electron sublevels. Notice also that the number of columns or families in each block exactly matches the number of electrons that occupy that sublevel. (This diagram also includes both the older and newer numbering schemes employed for the chemical families.)

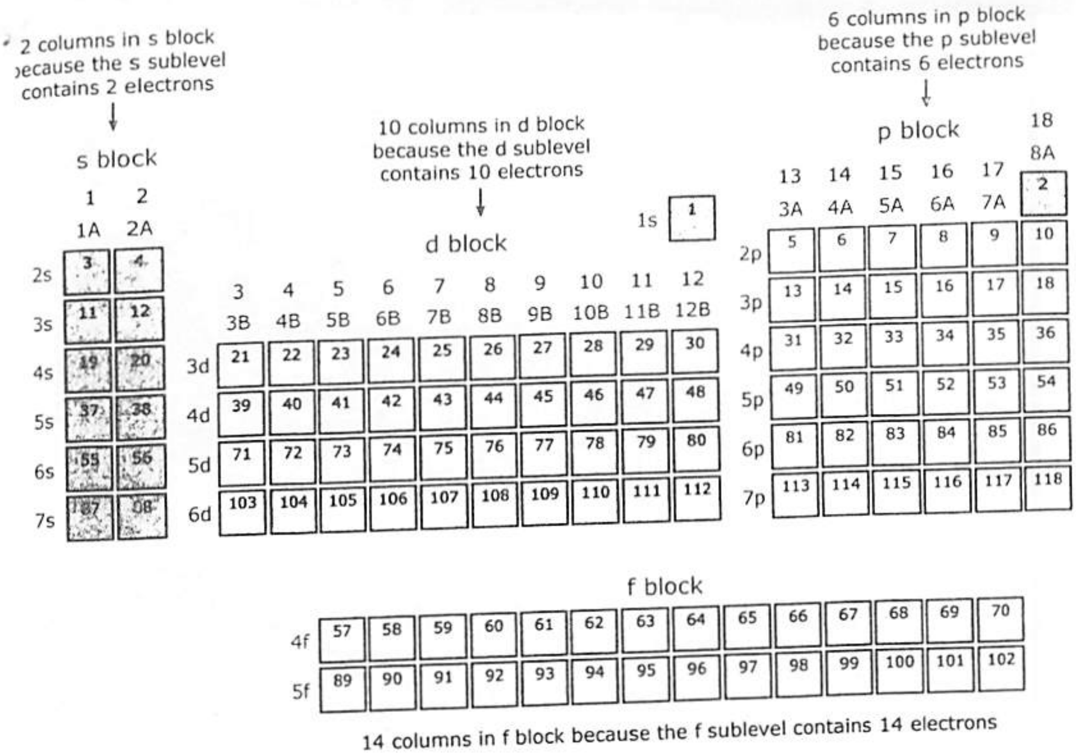


Figure 6.1.3 *The periodic table separated into its four main blocks*

Table 6.1.4 *Order of Filling Sublevels*

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

(2) Consider in Figure 6.1.3 the sublevels associated with each of the chemical periods in the table. Notice that as we "read" the sublevels from left to right like the words on a page, we follow exactly the order of filling sublevels specified by the Aufbau principle and summarized by the diagonal diagram in the previous chapter. That order is shown again in Table 6.1.4.

Note also that as we reach each noble gas at the end of a period, we move down and begin a new period. Each new period number specifies the value of the next energy level as we start filling the s orbital for elements on the left side of that period. Thus every new period represents a larger charge cloud or electron "shell" in which electrons spend most of their time farther from the nucleus.

(3) Finally, consider Figure 6.1.4, showing the extended form of the periodic table with the f block lanthanides and actinides inserted after barium and radium respectively. Note the consistent order and the obvious connection of electron configuration to element location in the periodic table.

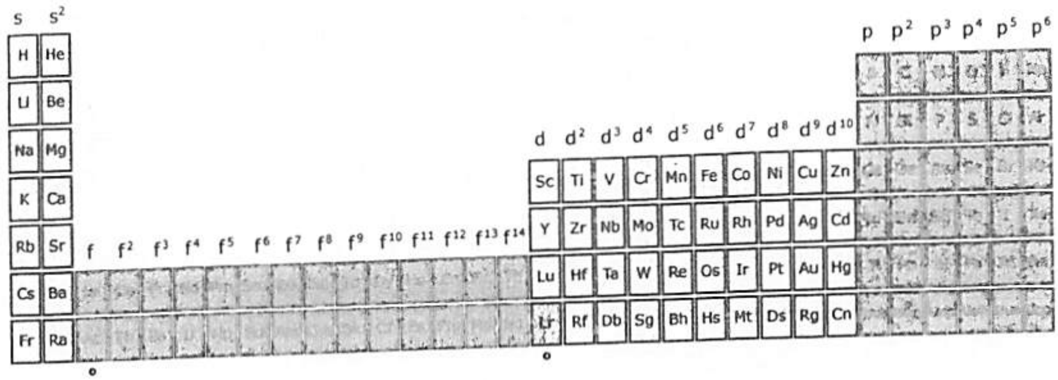


Figure 6.1.4 *Periodic table showing pattern of electron configuration*

Chemists, like all scientists, don't really believe in coincidences. We have discussed that the structure of the periodic table is based firmly on experimentation. The organization of the periods and chemical families results from and summarizes an enormous amount of empirical data gathered about the elements. It is truly amazing that this arrangement of the elements in the periodic table was essentially in place *before* the development of the quantum mechanical model of the atom! The fact that the electron configurations of the elements specified by quantum mechanics exactly matches their organization in the periodic table is one of the greatest successes in the history of science.

Far from being a coincidence, the exact correlation exists precisely because quantum mechanics provides the theoretical foundation for the experimentally derived periodic table. Stated another way, the quantum mechanical view of the atom is the "why" behind the "what" stated in the periodic law. Of course, it can also be said that the structure of the periodic table is arguably the best piece of supporting evidence for the quantum mechanical model of the atom.

Perhaps John Newlands wasn't far off the mark when he used a musical analogy in referring to elemental properties. Think of the millions of chemical reactions happening around us as symphonies played out by the instruments we call the chemical elements in nature's orchestral arrangement that is the periodic table.

6.1 Activity: The Elements of PowerPoint®

Question

How much information about an element can you present in a 10-minute presentation using Power Point® or another computer slide presentation software?

Procedure

1. Choose any one of the naturally occurring elements in the periodic table. Confirm the selection with your teacher to avoid duplicate selections among class members.
2. Using any and all resources available online or in print, produce a PowerPoint® presentation no longer than 10 minutes that includes as much information about your element as you can discover. Make sure to include the following: symbol, atomic number and mass, location on the periodic table, electron configuration, melting point, boiling point, common compounds, and uses. The presentation can include any videos showing reactions or compounds or any interesting properties of the element.

Results and Discussion

1. Be prepared to answer questions from your classmates and your teacher for a few minutes after your presentation.

6.1 Review Questions

1. What is the most important thing to know about the periodic table?

2. What significant modification to the periodic table occurred just before World War I?

3. Why do elements in the same chemical family have similar chemical properties?

4. (a) Where are the most metallic elements located on the periodic table?

   (b) Where are the most non-metallic elements located on the periodic table?

5. Consider the properties listed for eight different elements. Match each element on the left to its property on the right by writing the element's chemical symbol next to its property.

Element	Properties	Symbol
(a) chlorine	found in carbohydrates and an elemental gas in 21% of the atmosphere	
(b) silver	soft conductor that reacts explosively with water producing H_2 gas	
(c) neon	less than 30 g of this solid radioactive nonconductor exists on Earth	
(d) cesium	waxy yellow solid non-metal found in match heads, fertilizers, and detergents	
(e) oxygen	blue-gray metalloid used extensively in the computer industry	
(f) phosphorus	very reactive green gas used in the trenches in World War I	
(g) silicon	shiny solid that is the best conductor of heat and electricity	
(h) astatine	invisible unreactive gas used in lasers and some electric street signs	

6. Six different elemental properties are listed below corresponding to family numbers 1, 2, 6, 14, 17, and 18 in the periodic table. Write the appropriate family number next to each of the properties listed below. Each family number may be used only once.

Element Properties	Family Number
unreactive gas used in electric street signs and comprising 0.93% of the atmosphere	
shiny multivalent solid, good conductor, forms coloured compounds	
soft silvery solid, good conductor, reacts vigorously with water	
gray-white metalloid predicted by Mendeleev and discovered in 1886	
reactive metal present in bones and teeth possessing two valence electrons	
yellow-green gaseous non-metal and the most reactive of all the elements	

7. State four properties of elements classified as metals.

8. State four properties of elements classified as non-metals.

9. State a general rule for predicting the ion charges of many of the representative or main group elements.

10. Use your answer to question 8 above and write the formulas for the stable ions of the following:

(a) Be ____________________

(b) Te ____________________

(c) Cs ____________________

(d) Ra ____________________

(e) Ga ____________________

(f) Se ____________________

(g) In ____________________

11. Identify three properties of the elements belonging to each of the following chemical families:

(a) alkali metals

(b) alkaline earth metals

(c) halogens

(d) noble gases

12. Complete the following table by writing in the missing electron configurations. Highlight the outermost electrons in your answers. (Completed answers show outermost electrons in bold type.)

Group 2	Core Notation	Group 17	Core Notation	Group 18	Core Notation
Be		F	[He]$\mathbf{2s^2\,2p^5}$	He	
Mg		Cl		Ne	
Ca	[Ar]$\mathbf{4s^2}$	Br		Ar	
Sr		I		Kr	[Ar]$\mathbf{4s^2}\,3d^{10}\,\mathbf{4p^6}$
Ba	[Xe]$\mathbf{6s^2}$	At	[Xe] $\mathbf{6s^2}\,4f^{14}5d^{10}\mathbf{6p^5}$	Xe	
Ra				Rn	[Xe]$\mathbf{6s^2}\,4f^{14}5d^{10}\mathbf{6p^6}$