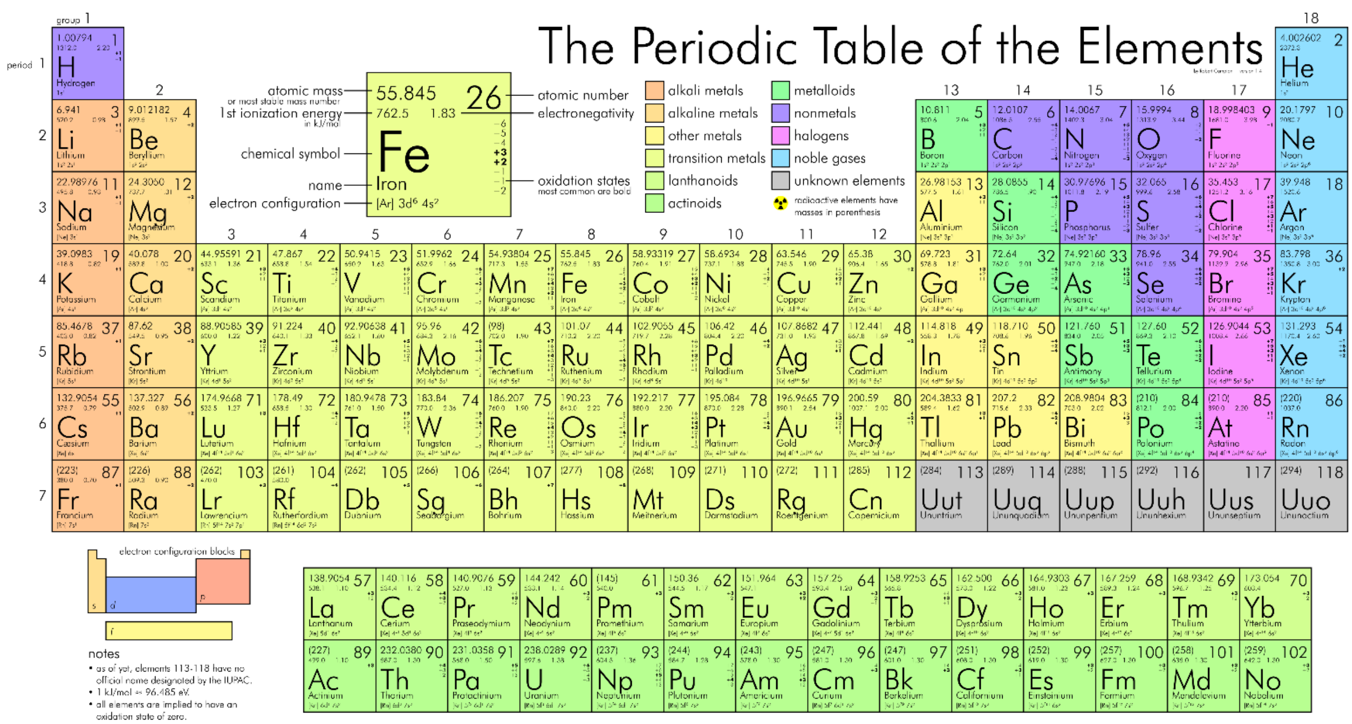
NTI Day 5 Lesson
Content Standard: Periodic Trends
Class: Chemistry **Teacher: K. Kelly**

The Periodic Table

*The **periodic table** is a chart displaying information about the elements. Elements are arranged in the table in a specific pattern that helps to predict their properties and to show their similarities and differences.*

Overview

The periodic table was developed by **Dmitri Mendeleev** in 1869. It provides a powerful tool for studying the elements and how they combine.



There are over 100 known elements, so it is necessary to use a systematic method to organize them. The periodic table indicates each element's atomic symbol, atomic number, and average atomic mass (also called atomic weight).

The placement of an element on the periodic table gives clues about the element's chemical and physical properties, including its melting point, density, hardness, and thermal and electrical conductivities.

Periods

The periodic table is so named because it is organized into "periods." A **period** is defined as *an interval required for a cycle to repeat itself*. In the periodic table, the periods are the horizontal rows that extend from left to right. These periods consist of as few as two elements and as many as thirty-two elements.

The repeating pattern across the periods of the periodic table is the filling of each energy level with electrons. For example, since hydrogen and helium are in period 1, they have electrons in only the 1st energy level. Since potassium, calcium, and bromine are in the 4th row or period, their outermost electrons are in the fourth energy level.

Electrons that are in the outermost energy level of an atom are called **valence electrons**. The position of an element within a period can help to determine the element's electron configuration and number of valence electrons.

Groups and Families

The division of elements into vertical **groups** by column creates **families** of elements. Elements in the same group all have the same number of valence electrons and therefore similar chemical properties.

Valence electrons help in predicting chemical reactions. For example, since hydrogen (H) is in group 1, it has 1 electron in its outer energy level. Chlorine (Cl), which is in group 17, has 7 electrons in its outer energy level and is one electron short of having a "full" outermost energy level. Thus, these two elements will readily combine to form the simple compound hydrochloric acid (HCl).

The **main group elements** are the elements located in groups 1, 2, and 13-18. These elements have some characteristic trends that are not followed by the transition metals in groups 3-12.

Some groups (or families) in the periodic table have special names. The properties of these groups are described below:

- **Group 1: Alkali metals**—Except for hydrogen, all of the elements in group 1 of the periodic table are alkali metals. They are soft metallic solids with low melting points, and they are the most reactive metals. Each element has only one electron in its outer shell, which explains why these elements tend to form ions with a charge of 1+.
- **Group 2: Alkaline earth metals**—All of the elements in group 2 of the periodic table are alkaline earth metals. They have higher melting points than alkali metals. Each element has only two electrons in its outer shell. Though these elements are also highly reactive, they are less reactive than alkali metals. An element in group 2 can have a full outer shell by losing two electrons, so group 2 elements tend to form ions that have 2+ charges.
- **Group 17: Halogens**—All of the elements in group 17 are halogens. They have low boiling points and low melting points. Each element in group 17 tends to gain an electron in its outer shell through bonding or forming an ion with a charge of 1-.
- **Group 18: Noble gases**—All of the elements in group 18 are noble gases. Each of these elements has a full outer shell of electrons and tends to be stable and unreactive. In general, noble gases do not react or combine with any element.
- **Groups 3-12: Transition metals**—Elements located in groups 3-12 on the periodic table are known as transition elements. These elements tend to have high heat and electrical conductivities. Some form 1+ ions, but some may form 2+ and 3+ ions as well.

Trends & the Periodic Table

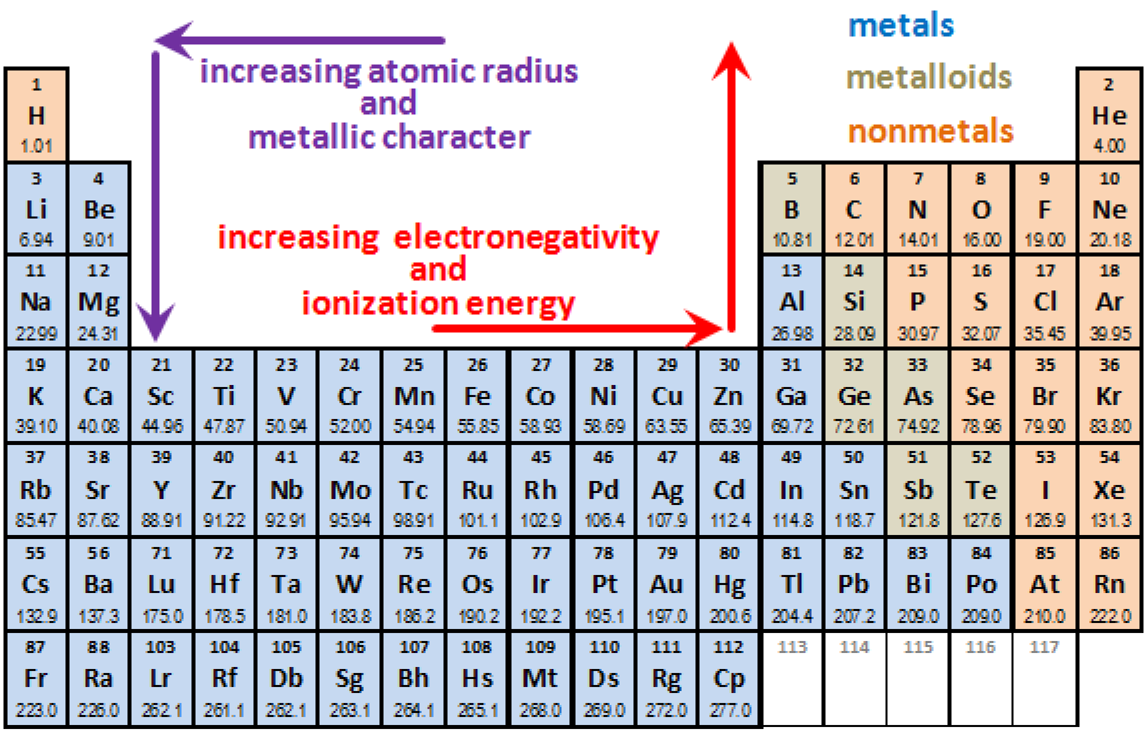
*There are many patterns and trends among the elements in the periodic table, including the **atomic radius**, **ionic radius**, **ionization energy**, and **electronegativity** of each element. Two main classes of elements are **metals** and **nonmetals**.*

Metals & Nonmetals

Atoms are arranged in the periodic table of the elements according to their properties. For example, metals are grouped together, and nonmetals (except for hydrogen) are grouped together. Elements have greater metallic character moving from right to left and from top to bottom within the periodic table. The stairstep line that begins between boron (B) and aluminum (Al) and moves down and right to polonium (Po) and astatine (At) is the dividing line between metals and nonmetals. The elements bordering the stairstep line have properties between those of metals and nonmetals, so these elements are called metalloids. This division is shown by the different colors in the periodic table below.

- **Metals** are the elements to the left of the stairstep. Metals are typically dense solids with a shiny luster. They tend to form **cations** (positively charged ions) and are capable of conducting electricity. Metals most often form ionic bonds with nonmetals. Metals, including alloys, are held together by metallic bonds between the atoms.

- **Nonmetals** include hydrogen and all of the elements to the right of the stairstep. They tend to have low densities, low melting points, and they do not conduct electricity. As solids, nonmetals have a dull luster. Nonmetals tend to gain electrons to become **anions** (negatively charged ions) when they form ionic bonds with metals, but they tend to form covalent bonds with other nonmetals.



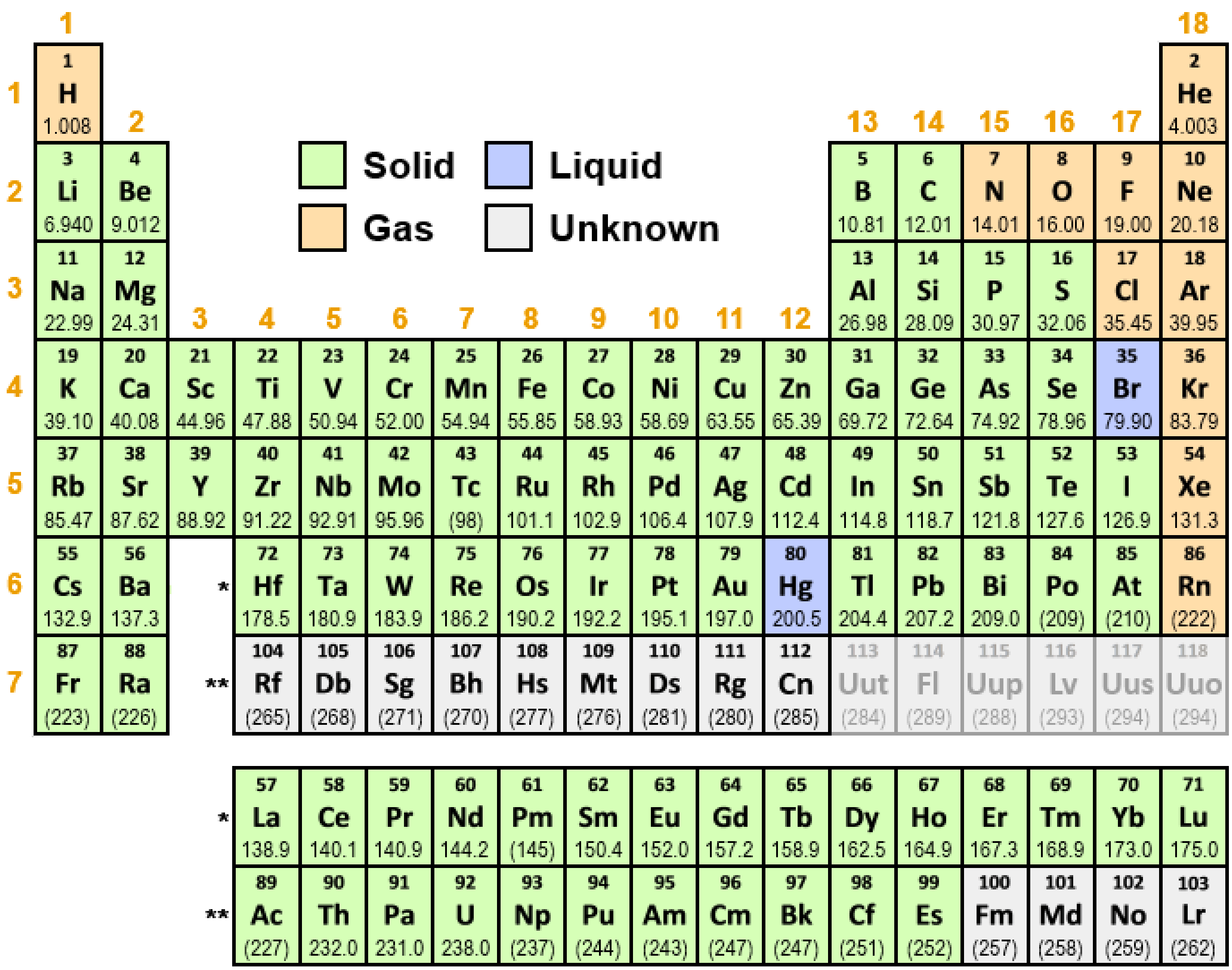
Trends in the Periodic Table

The arrangement of the elements in the periodic table allows certain trends in the physical and chemical properties of the elements to be predicted based on where they are located. These trends move left and right across the periods or up and down the groups. Some of these trends are described below:

- **Atomic radius** refers to the size of a neutral atom. Atomic radii generally increase in size from right to left across a period and from top to bottom down a group.

- **Electronegativity** refers to how strongly an atom attracts electrons within a bond, and it is an indicator of how reactive an element is. Electronegativity increases from left to right across a period and from bottom to top up a group. ***Fluorine*** has the greatest electronegativity and is the most reactive nonmetal. ***Francium*** is the most reactive metal because it has the lowest electronegativity.

- The **phase** of each element at room temperature is also related to its position on the periodic table. The elements on the left tend to exist as solids at room temperature. All of the elements that are gases at standard temperature and pressure are non-metals. Only two of the known elements are liquids at standard temperature and pressure. These two elements are mercury (Hg) and bromine (Br). Some of the elements on the periodic table have an unknown state of matter because they are unstable and do not occur in nature.



- The **number of valence electrons** present in a neutral atom is determined by the element's group (column) on the periodic table. Every element in a given group has the same number of valence electrons. For elements in the main groups (1-2, and 13-18), the number of valence electrons increases when moving from left to right along any period (row). The noble gases contain a full valence electron shell, which consists of eight electrons. The elements of other groups commonly form ions that enable them to fill their valence electron shells. Chlorine, for example, commonly gains an electron to fill its valence shell, making it an ion with a charge of 1-. Sodium, at the other end of the periodic table, commonly loses one electron, which allows it to fill its valence shell by becoming an ion of charge 1+.

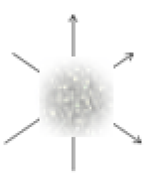
Electron Configurations

The electron configuration of an element shows how the electrons are arranged around the nucleus of each atom when the atom is in its lowest energy state.

Electron Shells, Subshells & Atomic Orbitals

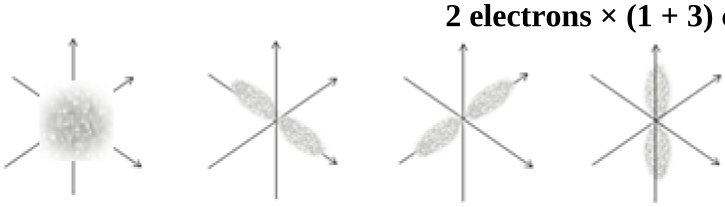
The arrangement of electrons around the nucleus of an atom can be described in terms of *energy levels*, *subshells*, and *atomic orbitals*.
The highest **energy level** number (1 through 7) for the electrons in an atom corresponds to the period (or row) in the periodic table to which that atom belongs. Because there are 7 periods in the table, there are 7 energy levels. For example, hydrogen (H) is in the first period, so it has only one energy level. Sodium is in the third period, so it has three energy levels. As the number describing the energy level increases, the energy of that level also increases.
Next, each energy level has *sublevels*, or **subshells**, which contain different **atomic orbitals**. There are four main types of subshells (s, p, d, and f), which each contain atomic orbitals. The orbitals have different shapes and often different orientations in space.

Each type of orbital can accommodate a different number of electrons. According to the electron cloud model, the first energy level (numbered 1) only has one s orbital. Therefore, it can only hold a maximum of two electrons. S orbitals are approximately spherical. A model of an s orbital is shown below.



The next energy level (numbered 2) has one s orbital and three p orbitals. P orbitals are shaped approximately like 3-dimensional figure eights, and they are oriented in space differently. Each p orbital can hold two electrons. Therefore, the second energy level can hold a maximum of eight electrons.

2 electrons × (1 + 3) clouds = 8 electrons



The shape of an atomic orbital is defined through the use of mathematics. It describes regions in space where electrons are most likely to be located.

The orbitals, their subshells, the number of orbitals, and the number of electrons each subshell can hold are shown in the table below. Note that the d and f subshells are shown for completion only; they are out of the scope of this lesson.

Energy Level	Subshells	# of orbitals	# of electrons it can hold
1	1s	1	2
2	2s	1	2
	2p	3	6
3	3s	1	2
	3p	3	6
	3d	5	10
4	4s	1	2
	4p	3	6
	4d	5	10
	4f	7	14

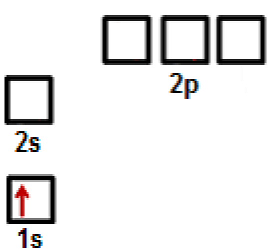
Since there are often multiple sublevels in an energy level, the total number of electrons in an energy level varies. The first energy level can hold up to 2 electrons, the second energy level can hold up to 8 electrons, the third energy level can hold up to 18 electrons, and the fourth energy level can hold up to 32 electrons.

The Aufbau Principle

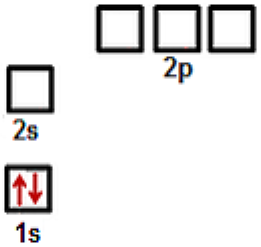
As mentioned above, each energy level or electron shell often contains multiple subshells. For example, the 3s, 3p, and 3d subshells all belong to the third electron shell. However, despite being part of the same energy level, they do not possess the same amount of energy; the 3s orbital is lower in energy than the 3p orbitals, but the 3p orbitals all have the same energy. Subshells with lower energy fill first according to an accepted rule called the *Aufbau principle*. The approximate relative energies of the subshells for each energy level are indicated in the diagram below.

Each box in the diagram represents one atomic orbital. Each row of boxes represents one subshell. According to the Aufbau principle, electrons position themselves so that they fill the orbitals with the lowest energy first. So, the 1s orbital is always filled before the 2s orbital, the 2s orbital is always filled before the 2p orbital, and so on. This information can be used to fill in orbital diagrams.

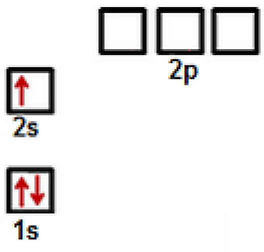
Since each orbital can carry two electrons, and the two electrons are paired—one spin up and one spin down—two arrows are usually placed in each box in an orbital diagram—one pointing up and one pointing down. Each arrow represents one electron. As electrons fill in the lowest-energy orbitals first, electrons should be added to the lowest boxes first.



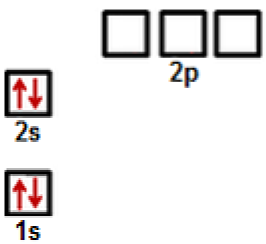
One Electron



Two Electrons



Three Electrons



Four Electrons

Writing Electron Configurations

Electron configurations are generally written in the following format.

$$nx^y$$

where n is the energy level number, x describes the shape of the atomic orbital, and y designates how many electrons are in the subshell. However, since electrons fill the orbitals with the lowest energy first, electron shells do not fill completely in sequence. They fill the energy levels in the order:

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

Finally, since the electron configurations of elements can be quite long, it is possible to write a shorthand version. To do this, use a periodic table to find the noble gas that immediately precedes the element of interest, write the chemical symbol of this noble gas in brackets, and then only write the portion of the electron configuration that follows the configuration of the noble gas. For example, the shorthand version of the electron configuration for cesium is:

$$[Xe]\ 6s^1$$

The Organization of the Periodic Table & Orbital Blocks

The electron configurations of elements may be better understood by dividing the periodic table into s orbital, p orbital, d orbital, and f orbital blocks. The width of each orbital block is related to the maximum number of electrons that can be held by a particular orbital (i.e., s orbitals can hold up to two electrons, so the s block is two elements wide; p orbitals can hold up to six electrons, so the p block is six elements wide, etc.).

$1s^1$																	$1s^2$
$2s^1$	$2s^2$											$2p^1$	$2p^2$	$2p^3$	$2p^4$	$2p^5$	$2p^6$
$3s^1$	$3s^2$											$3p^1$	$3p^2$	$3p^3$	$3p^4$	$3p^5$	$3p^6$
$4s^1$	$4s^2$	$3d^1$	$3d^2$	$3d^3$	$3d^4$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^9$	$3d^{10}$	$4p^1$	$4p^2$	$4p^3$	$4p^4$	$4p^5$	$4p^6$
$5s^1$	$5s^2$	$4d^1$	$4d^2$	$4d^3$	$4d^4$	$4d^5$	$4d^6$	$4d^7$	$4d^8$	$4d^9$	$4d^{10}$	$5p^1$	$5p^2$	$5p^3$	$5p^4$	$5p^5$	$5p^6$
$6s^1$	$6s^2$	$5d^1$	$5d^2$	$5d^3$	$5d^4$	$5d^5$	$5d^6$	$5d^7$	$5d^8$	$5d^9$	$5d^{10}$	$6p^1$	$6p^2$	$6p^3$	$6p^4$	$6p^5$	$6p^6$
$7s^1$	$7s^2$	$6d^1$	$6d^2$	$6d^3$	$6d^4$	$6d^5$	$6d^6$	$6d^7$	$6d^8$	$6d^9$	$6d^{10}$						
		$4f^1$	$4f^2$	$4f^3$	$4f^4$	$4f^5$	$4f^6$	$4f^7$	$4f^8$	$4f^9$	$4f^{10}$	$4f^{11}$	$4f^{12}$	$4f^{13}$	$4f^{14}$		
		$5f^1$	$5f^2$	$5f^3$	$5f^4$	$5f^5$	$5f^6$	$5f^7$	$5f^8$	$5f^9$	$5f^{10}$	$5f^{11}$	$5f^{12}$	$5f^{13}$	$5f^{14}$		

Again, note that blocks d and f are out of the scope of this lesson.

Tips for Writing Electron Configurations

With the help of a periodic table, neutral atoms of elements can be easily identified from their electron configurations.

1 H 1.01																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.91	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	71 Lu 175.0	72 Hf 178.5	73 Ta 181.0	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po 209.0	85 At 210.0	86 Rn 222.0
87 Fr 223.0	88 Ra 226.0	103 Lr 262.1	104 Rf 261.1	105 Db 262.1	106 Sg 263.1	107 Bh 264.1	108 Hs 265.1	109 Mt 268.0	110 Ds 269.0	111 Rg 272.0	112 Cp 277.0	113	114	115	116	117	

57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm 144.9	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0
89 Ac 227.0	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np 237.1	94 Pu 244.1	95 Am 243.1	96 Cm 247.1	97 Bk 247.1	98 Cf 251.1	99 Es 252.1	100 Fm 257.1	101 Md 258.1	102 No 259.1

For example, if given the following electron configuration:

$$1s^2 2s^2 2p^6 3s^2 3p^5$$

the element may be identified in one of two ways.

- Add the superscripts together. The sum of the superscripts will be equal to the total number of electrons in the atom. If the atom is neutral, then the number of electrons is equal to the number of protons, or the atomic number of the atom. Using this method, there are 17 electrons (2 + 2 + 6 + 2 + 5), so the element is chlorine.
- Look at the last expression in the electron configuration. In this case, the last expression is $3p^5$. Using the two tables given above, chlorine corresponds to the $3p^5$ block