Stockton Unified School District
Since 1852

# Chemistry
# 3-30 thru 4-3
# 2020

# Chemistry Daily Lesson Log

## Week of 3/30/2020 - 4/3/2020

Each day

1. Check out the lesson for the day
2. Complete the activities listed
3. Complete a learning log entry for the day. The learning log must be completed each day (even if you don't finish all the work, still state what you learned), make sure to write the date and activities completed, and what you learned from them.

| Lesson # | Materials | Due @ the end of the lesson - check each item off as you complete it! |
|---|---|---|
| 6<br>(Monday 3-30) | ❏ Periodic Table Coloring Activity<br>❏ Colored Pencils | ❏ Color the periodic table according to the directions<br>❏ Answer the questions on the final page of the worksheet entitled "Family Ties"<br>❏ Learning Log |
| 7<br>(Tuesday 3-31) | ❏ Stemscopedia on the periodic table pages 6-9 ONLY | ❏ Write a summary of what you learned<br>❏ Write a list of questions you still have about how the periodic table operates<br>❏ Learning Log |
| 8<br>(Wednesday 4-1) | ❏ POGIL on Coulombic Attraction | ❏ Complete POGIL activity and answer questions as you go<br>❏ Learning Log |
| 9<br>(Thursday 4-2) | ❏ POGIL on Periodic Trends | ❏ Complete POGIL activity and answer questions as you go<br>❏ Learning Log |
| 10<br>(Friday 4-3) | ❏ Stemscopedia pages 9-12 | ❏ Write a summary of what you learned<br>❏ Write a list of questions you still have about how the periodic table operates<br>❏ Learning Log |

**Below are some video resources that you may find helpful. They aren't required, but you may find them useful!**

Coulomb's Law simulation: https://phet.colorado.edu/sims/html/coulombs-law/latest/coulombs-law_en.html

Periodic Table: https://www.youtube.com/watch?v=tc9tEUqUmKw

Periodic Trends: https://www.youtube.com/watch?v=XK-WTYncldA

Name: Period:

Chemistry Learning Log - Complete one row EVERY DAY

| | |
|---|---|
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) **I learned that...** |
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) **I learned that...** |
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) **I learned that...** |
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) I **learned that...** |
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) **I learned that...** |

Name: Period:

Learning Log

| | |
|---|---|
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) I learned that... |
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) I learned that... |
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) I learned that... |
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) I learned that... |
| Date:<br><br>What did you do today?<br>(Bullet points okay) | What did you learn from these activities? (3 - 5 sentences) I learned that... |

Name: ________________________________________ Period: _______

# Color Coding the Periodic Table

There are a number of major groups with similar properties. They are as follows:

Hydrogen: This element does not match the properties of any other group so it stands alone. It is placed above group 1 but it is not part of that group. It is a very reactive, colorless, odorless gas at room temperature. (1 outer level electron)

Group 1: Alkali Metals – These metals are extremely reactive and are never found in nature in their pure form. They are silver colored and shiny. Their density is extremely low so that they are soft enough to be cut with a knife. (1 outer level electron)

Group 2: Alkaline-earth Metals – Slightly less reactive than alkali metals. They are silver colored and more dense than alkali metals. (2 outer level electrons)

Groups 3 – 12: Transition Metals – These metals have a moderate range of reactivity and a wide range of properties. In general, they are shiny and good conductors of heat and electricity. They also have higher densities and melting points than groups 1 & 2. (1 or 2 outer level electrons)

Lanthanides and Actinides: These are also transition metals that were taken out and placed at the bottom of the table so the table wouldn't be so wide. The elements in each of these two periods share many properties. The lanthanides are shiny and reactive. The actinides are *all* radioactive and are therefore unstable. Elements 95 through 103 do not exist in nature but have been manufactured in the lab.

Group 13: Boron Group – Contains one metalloid and 4 metals. Reactive. Aluminum is in this group. It is also the most abundant metal in the earth's crust. (3 outer level electrons)

Group 14: Carbon Group – Contains on nonmetal, two metalloids, and two metals. Varied reactivity. (4 outer level electrons)

Group 15: Nitrogen Group – Contains two nonmetals, two metalloids, and one metal. Varied reactivity. (5 outer level electrons)

Group 16: Oxygen Group – Contains three nonmetals, one metalloid, and one metal. Reactive group. (6 outer level electrons)

Groups 17: Halogens – All nonmetals. Very reactive. Poor conductors of heat and electricity. Tend to form salts with metals. Ex. NaCl: sodium chloride also known as "table salt". (7 outer level electrons)

Groups 18: Noble Gases – Unreactive nonmetals. All are colorless, odorless gases at room temperature. All found in earth's atmosphere in small amounts. (8 outer level electrons)

Name: ____________________________________________ Period: ________

***Color Coding the Periodic Table Student Worksheet***

This worksheet will help you understand how the periodic table is arranged. Your teacher will give you a copy of the periodic table to color. Using map pencils, color each group on the table as follows:

1. Number the top (columns) of the periodic table 1-18
2. Number the side (rows) of the period table 1-7
3. Color the square for Hydrogen pink.
4. Lightly color all metals yellow.
5. Place black dots in the squares of all alkali metals.
6. Draw a horizontal line across each box in the group of alkaline earth metals.
7. Draw a diagonal line across each box of all transition metals.
8. Color the metalloids purple.
9. Color the nonmetals orange.
10. Draw small brown circles in each box of the halogens.
11. Draw checkerboard lines through all the boxes of the noble gases.
12. Using a black color, trace the zigzag line that separates the metals from the nonmetals.
13. Color all the lanthanides red.
14. Color all the actinides green.

When you are finished, make sure you have a key that indicates which color identifies which group on your periodic table!

Name: ______________________________ Period: ________

# Periodic Table of the Elements

| 1 IA 1A | 2 IIA 2A | 3 IIIB 3B | 4 IVB 4B | 5 VB 5B | 6 VIB 6B | 7 VIIB 7B | 8 | 9 VIII 8 | 10 | 11 IB 1B | 12 IIB 2B | 13 IIIA 3A | 14 IVA 4A | 15 VA 5A | 16 VIA 6A | 17 VIIA 7A | 18 VIIIA 8A |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 1 **H** Hydrogen 1.008 | | | | | | | | | | | | | | | | | 2 **He** Helium 4.003 |
| 3 **Li** Lithium 6.941 | 4 **Be** Beryllium 9.012 | | | | | | | | | | | 5 **B** Boron 10.811 | 6 **C** Carbon 12.011 | 7 **N** Nitrogen 14.007 | 8 **O** Oxygen 15.999 | 9 **F** Fluorine 18.998 | 10 **Ne** Neon 20.180 |
| 11 **Na** Sodium 22.990 | 12 **Mg** Magnesium 24.305 | | | | | | | | | | | 13 **Al** Aluminum 26.982 | 14 **Si** Silicon 28.086 | 15 **P** Phosphorus 30.974 | 16 **S** Sulfur 32.066 | 17 **Cl** Chlorine 35.453 | 18 **Ar** Argon 39.948 |
| 19 **K** Potassium 39.098 | 20 **Ca** Calcium 40.078 | 21 **Sc** Scandium 44.956 | 22 **Ti** Titanium 47.867 | 23 **V** Vanadium 50.942 | 24 **Cr** Chromium 51.996 | 25 **Mn** Manganese 54.938 | 26 **Fe** Iron 55.845 | 27 **Co** Cobalt 58.933 | 28 **Ni** Nickel 58.693 | 29 **Cu** Copper 63.546 | 30 **Zn** Zinc 65.38 | 31 **Ga** Gallium 69.723 | 32 **Ge** Germanium 72.631 | 33 **As** Arsenic 74.922 | 34 **Se** Selenium 78.971 | 35 **Br** Bromine 79.904 | 36 **Kr** Krypton 83.798 |
| 37 **Rb** Rubidium 85.468 | 38 **Sr** Strontium 87.62 | 39 **Y** Yttrium 88.906 | 40 **Zr** Zirconium 91.224 | 41 **Nb** Niobium 92.906 | 42 **Mo** Molybdenum 95.95 | 43 **Tc** Technetium 98.907 | 44 **Ru** Ruthenium 101.07 | 45 **Rh** Rhodium 102.906 | 46 **Pd** Palladium 106.42 | 47 **Ag** Silver 107.868 | 48 **Cd** Cadmium 112.414 | 49 **In** Indium 114.818 | 50 **Sn** Tin 118.711 | 51 **Sb** Antimony 121.760 | 52 **Te** Tellurium 127.6 | 53 **I** Iodine 126.904 | 54 **Xe** Xenon 131.294 |
| 55 **Cs** Cesium 132.905 | 56 **Ba** Barium 137.328 | 57-71 | 72 **Hf** Hafnium 178.49 | 73 **Ta** Tantalum 180.948 | 74 **W** Tungsten 183.84 | 75 **Re** Rhenium 186.207 | 76 **Os** Osmium 190.23 | 77 **Ir** Iridium 192.217 | 78 **Pt** Platinum 195.085 | 79 **Au** Gold 196.967 | 80 **Hg** Mercury 200.592 | 81 **Tl** Thallium 204.383 | 82 **Pb** Lead 207.2 | 83 **Bi** Bismuth 208.980 | 84 **Po** Polonium [208.982] | 85 **At** Astatine 209.987 | 86 **Rn** Radon 222.018 |
| 87 **Fr** Francium 223.020 | 88 **Ra** Radium 226.025 | 89-103 | 104 **Rf** Rutherfordium [261] | 105 **Db** Dubnium [262] | 106 **Sg** Seaborgium [266] | 107 **Bh** Bohrium [264] | 108 **Hs** Hassium [269] | 109 **Mt** Meitnerium [278] | 110 **Ds** Darmstadtium [281] | 111 **Rg** Roentgenium [280] | 112 **Cn** Copernicium [285] | 113 **Nh** Nihonium [286] | 114 **Fl** Flerovium [289] | 115 **Mc** Moscovium [289] | 116 **Lv** Livermorium [293] | 117 **Ts** Tennessine [294] | 118 **Og** Oganesson [294] |

| | | | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| **Lanthanide Series** | 57 **La** Lanthanum 138.905 | 58 **Ce** Cerium 140.116 | 59 **Pr** Praseodymium 140.908 | 60 **Nd** Neodymium 144.243 | 61 **Pm** Promethium 144.913 | 62 **Sm** Samarium 150.36 | 63 **Eu** Europium 151.964 | 64 **Gd** Gadolinium 157.25 | 65 **Tb** Terbium 158.925 | 66 **Dy** Dysprosium 162.500 | 67 **Ho** Holmium 164.930 | 68 **Er** Erbium 167.259 | 69 **Tm** Thulium 168.934 | 70 **Yb** Ytterbium 173.055 | 71 **Lu** Lutetium 174.967 |
| **Actinide Series** | 89 **Ac** Actinium 227.028 | 90 **Th** Thorium 232.038 | 91 **Pa** Protactinium 231.036 | 92 **U** Uranium 238.029 | 93 **Np** Neptunium 237.048 | 94 **Pu** Plutonium 244.064 | 95 **Am** Americium 243.061 | 96 **Cm** Curium 247.070 | 97 **Bk** Berkelium 247.070 | 98 **Cf** Californium 251.080 | 99 **Es** Einsteinium [254] | 100 **Fm** Fermium 257.095 | 101 **Md** Mendelevium 258.1 | 102 **No** Nobelium 259.101 | 103 **Lr** Lawrencium [262] |

Name: ________________________________________ Period: _______

# Family Ties

Student Worksheet

Follow the instructions below to label the major groups and divisions of the periodic table.

1. The vertical columns on the periodic table are called ____________.
2. The horizontal rows on the periodic table are called _____________.
3. Most of the elements in the periodic table are classified as _____________.
4. The elements that touch the zigzag line are classified as _______________.
5. The elements in the far upper right corner are classified as______________.
6. Elements in the first group have one outer shell electron and are extremely reactive. They are called ___________ ______________.
7. Elements in the second group have 2 outer shell electrons and are also very reactive. They are called ______________ ______________ ________________.
8. Elements in groups 3 through 12 have many useful properties and are called _________________ _______________.
9. Elements in group 17 are known as "salt formers". They are called _________________.
10. Elements in group 18 are very unreactive. They are said to be "inert". We call these the ______________ ______________.
11. The elements at the bottom of the table were pulled out to keep the table from becoming too long. The first period at the bottom called the _________________.
12. The second period at the bottom of the table is called the _____________________.

**Objectives**

- Describe the structure of an atom.
- Use the periodic table as a model to predict the relative properties of elements.
- Use the periodic table to predict trends in size, reactivity, and electronegativity.
- Distinguish between ionic, covalent, and metallic bonds.

**Key Terms**

Atoms
Electron
Protons
Nucleus
Neutrons
Isotope
Energy levels
Valence electrons
Groups
Periods
Octet
Ions
Ionic bonding
Covalent bonding
Metallic bonding
Electronegativity
Atomic radii

# Periodic Table and Element Structure

Everything in the universe, including all matter on Earth, is formed from unique combinations of tiny, neutral particles called **atoms**. These particles are so small that they cannot be seen with even the best microscope. Atoms of the same kind come together to form elements. To this date, there are 118 elements that have been discovered. The elements' names have been placed in a chart called the periodic table. The placement of each element in this "warehouse" is based on characteristics of that particular atom.

## The Atom

Even though atoms cannot be seen by the naked eye, the concept has been in existence for over 2,000 years. Around 442 BC, the Greek theorist Democritus proposed a theory on the premise that if a stone were continuously cut in half, you would eventually come to a point at which the stone would be too small to halve. He called this *atomos*, which translates to mean "indivisible." An atom is defined as the smallest component of an element that has the properties of that element.

Later theories led the way to the discovery of subatomic particles, or substructures. Today we understand that the atom consists of **electrons**, **protons**, and **neutrons**. Each subparticle was discovered by different investigations.

*Electrons*

Electrons were discovered in 1897 by the British physicist J. J. Thomson in his cathode ray experiment. Thomson discovered the negatively charged particle that orbits the atom's center by sending an electric current through an empty glass cathode tube. The rays were bent toward the positively charged plate and away from the negatively charged plate. This led to the theory that atoms consisted of a very small particle with a negative charge, called the electron. An electron is defined as a negatively charged subatomic particle with no known substructure.

*Protons*

A proton is the positively charged particle found in the center of the atom. Protons were discovered in 1911 by Ernest Rutherford, using the gold foil experiment. This experiment shot a beam of alpha particles at a thin, gold foil sheet. His theory was that particles should pass through the gold foil. However, a small percentage of the atoms was deflected. Figure 1 shows how the gold foil experiment was set up.

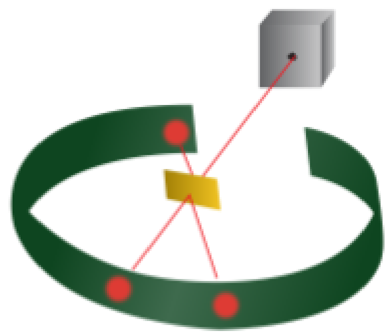


Figure 1

This led to the theory that a very small region of an atom has a positive charge, leading to the subparticle called a proton. He further theorized that these protons were in a dense region of the atom he termed the **nucleus**.

*Neutrons*

In 1932, James Chadwick, an English physicist who worked closely with Rutherford, discovered neutrons. He based his findings on his and others' work that measured the mass of atoms; yet when compared to the number of protons, those numbers were different. He proposed that a neutron was another particle in the nucleus that had a mass similar to the proton but no charge.

## The Nature of Science in Action

**Scientific Investigations Use a Variety of Methods**

Scientific inquiry is characterized by a common set of values that include logical thinking, precision, open-mindedness, objectivity, skepticism, replicability of results, and honest and ethical reporting of findings.

The inquiry of these scientists set the stage for further investigation of the atom, including structure and energy resources.

*Isotopes*

There are elements with atoms that are not completely identical. These atoms have the same number of electrons and protons and, therefore, possess the same chemical properties, but the mass of the atoms differs. This is due to the number of neutrons found in the atoms' nuclei. These are called **isotopes**, which are atoms of the same element that differ in number of neutrons. Figure 2 shows the two naturally occurring isotopes of helium. Not all elements have known isotopes. Those that do have isotopes can have as few as one, such as lithium, $^{6}Li$ and $^{7}Li$ (stated as lithium 6 and lithium 7). Tin currently has the most known isotopes, with 10.

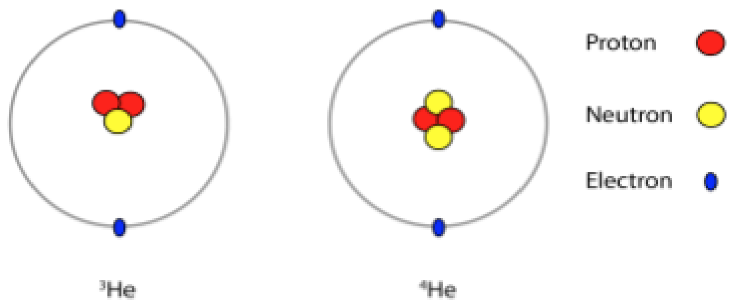



**Figure 2**

## Atomic Structure

Once scientists understood the parts of an atom, it was only a matter of time before there was inquiry into the movement, shape, and structure of its parts. During the late 19th century, scientists began to theorize about the shape and structure of the atom.

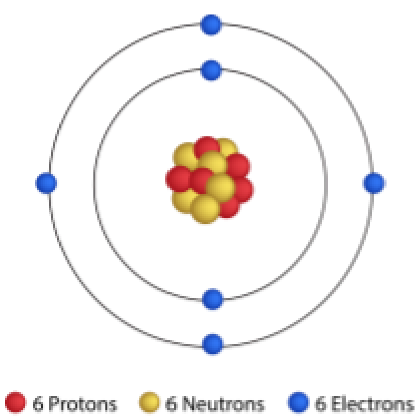



Figure 3

*The Bohr Model*

Niels Bohr's model was developed in 1913 with concentric circles around a central nucleus. He proposed that these concentric circles would have set energies and electrons would travel in those orbits, or **energy levels**, much like planets orbiting the Sun. An energy level was defined as a discrete region around the nucleus where electrons can exist. A diagram of a carbon atom Bohr model can be seen in figure 3. Electrons were able to move from one energy level to another if the atom was exposed to enough energy to allow for the movement. He also hypothesized that as these electrons moved to a lower energy level, that energy was emitted in the form of a photon, thus emitting light.

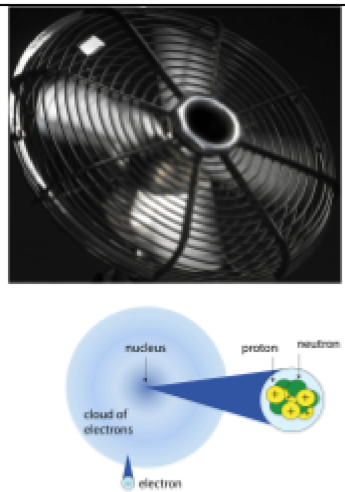



Figure 4

Although Bohr's model seemed to answer most questions about atomic behavior for hydrogen, there were still questions about atoms with more than one electron. Through the work of Louis de Broglie, Werner Heisenberg, and Erwin Schrödinger, the quantum model of the atom was established. The quantum model states that there are regions within the atom where there is a high probability of finding the electrons. These regions are referred to as electron orbitals. The electron cloud is the region where there is the highest probability of finding an electron within an orbital. Electrons do not follow a set path, like the planets do, but instead move within the area of the electron cloud at extremely high speeds. A good illustration is a fan's moving blades (figure 4). When turned on, the moving blades look like one continuous blade instead of the individual ones. You know the blades are moving within that area; however, you cannot say for certain where a specific blade is at any precise moment. This is how you can describe electrons and their movement in the electron cloud. We know they travel in a confined area (the electron cloud), but we cannot say for certain where each one is found at any specific time. This is known as the Heisenberg uncertainty principle.

The accepted modern atomic theory states that an atom consists of a positively charged, centrally located nucleus containing protons and neutrons. Negatively charged electrons are found in the area around the nucleus in the electron cloud. Protons are used to determine the atom's identity, while the **valence electrons** will determine the atom's reactivity. Valence electrons are found in the outermost energy level of an atom.

## Organization of the Elements

Figure 5

*Dmitri Mendeleev*

As atoms of elements were being studied in the late nineteenth century, several scientists were looking for a way to organize the elements to keep track of known elements and add new ones as they were discovered. The person given credit for developing this organization was a scientist and teacher named Dmitri Mendeleev (figure 5). In 1869, Mendeleev began by writing the physical and chemical properties of each element on a card. When he organized them by atomic mass, he began to see a pattern. Certain characteristics occurred periodically, or in a predictable manner. As he continued to place the 60 elements that were known at the time, he realized there were some open spaces on the table. He predicted that elements would be discovered that would fill those spots, and he even went so far as to describe what properties those elements would have. When gallium (1875) and germanium (1886) were discovered and their properties were a close match to what Mendeleev predicted, the scientific community realized the organization of the table was a powerful tool.

*The Modern Periodic Table*

So, does the periodic table today look like the one Mendeleev made in 1869? Although Mendeleev's organization made sense, there were a few elements that seemed out of place. In 1913, British scientist Henry Moseley used X-rays to investigate properties of the elements. He found that as you moved up one element on Mendeleev's chart of elements, there was a slight change in the wavelength. He thus proposed an increase in charge. This would correlate to the number of protons in each atom. Later that year, the term *atomic number* was coined and designated to represent the number of protons found in an atom.

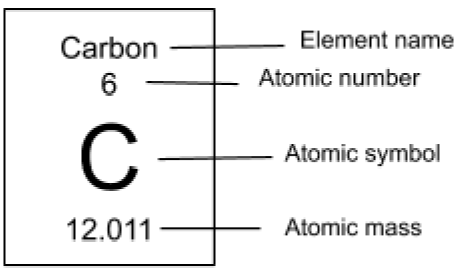



Figure 6

When organizing the elements using the the atomic number, the anomalies of reactivity in Mendeleev's table were resolved. The modern periodic table has been organized by atomic number since then. Figure 6 shows how elements are represented with atomic numbers on the periodic table.

*Periodic Law*

As elements were arranged by atomic number in the rows and columns of the table, a pattern of properties emerged. The columns of the table were referred to as **groups**, and the elements in those groups had similar chemical properties. The rows of the table were referred to as **periods**. The repeating pattern of properties as you go across a period is a reflection of the periodic law. The periodic law states that the physical and chemical properties of the elements are periodic functions of their atomic numbers.

## The Periodic Table

As the elements were organized by atomic number, the patterns of behavior became very obvious. That allowed for some generalized classifications of the elements. As seen in figure 7, there are three main classifications: metals (blue), nonmetals (white), and metalloids (yellow). It is very clear that most elements on the periodic table are metals.

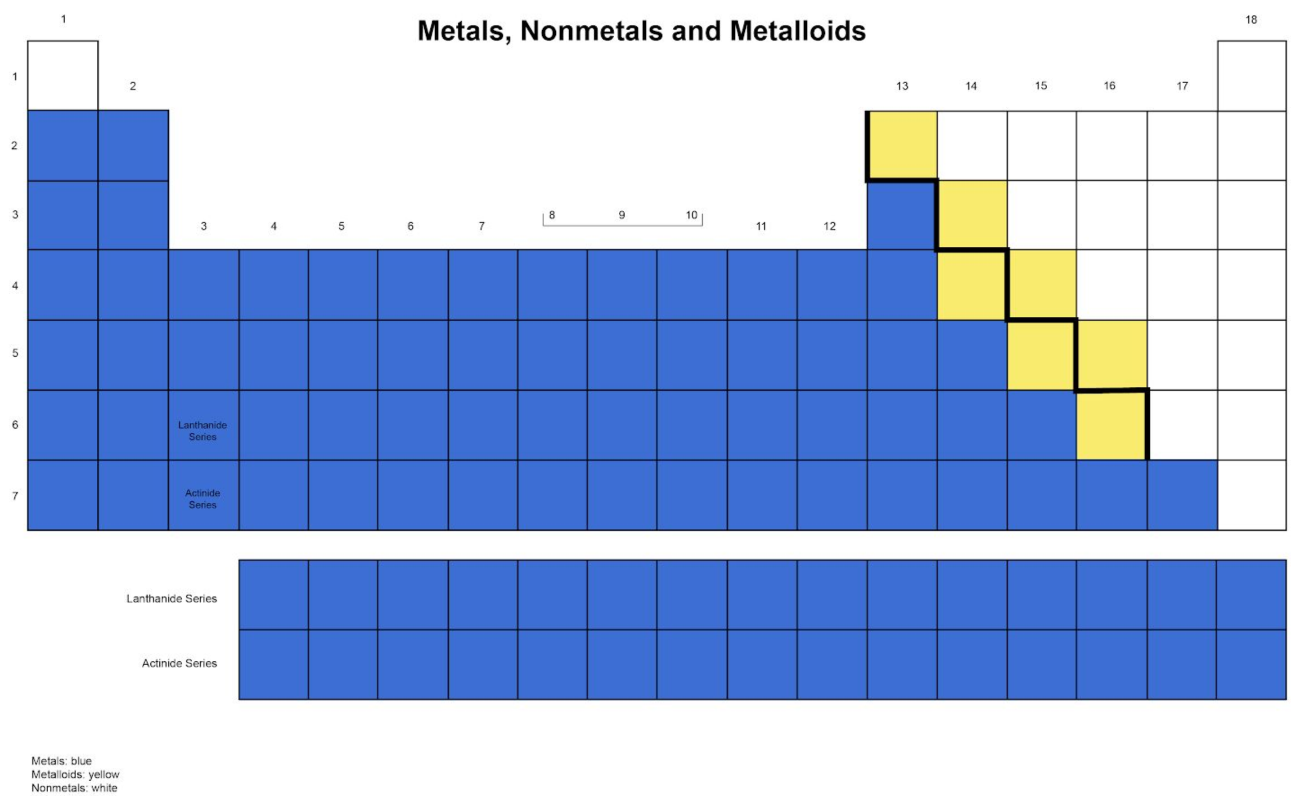


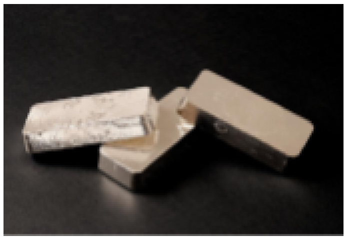

**Figure 7**

### The Nature of Science in Action

**Scientific Knowledge Assumes an Order and Consistency in Natural Systems**

Scientific knowledge is based on the assumption that natural laws operate today as they did in the past and they will continue to do so in the future.

Periodic law continues to be proven as more elements are discovered and added to the periodic table.



Silver Bars, a metal

Sulfur, a nonmetal

Silicon (crystallized), a metalloid

Metals are located to the left of the stairstep of the periodic table and are shown in blue in figure 7. Most metals share similar properties; they are generally dense, shiny solids (except one) with high melting points. Metals are malleable and ductile, meaning they can be hammered thin or pulled into thin wires without breaking. Metals are also good conductors of both heat and electricity.

Nonmetals may be found to the right side of the periodic table and are shown in white in figure 7. They are very different from metals. The surface of most nonmetals is dull, and they are poor conductors of both heat and electricity. Most nonmetals will melt at low temperatures, and they also have a low density. Unlike metals, nonmetals are brittle and will break when stretched or hammered.

Metalloids are elements that have properties of both metals and nonmetals and are located in the stairstep area between the metals and nonmetals. They are shown in yellow in figure 7. Like metals, metalloids may be shiny. However, metalloids may also be dull, like many nonmetals. They can conduct both heat and electricity, but not as well as metals.

*Valence Electrons and Groups*

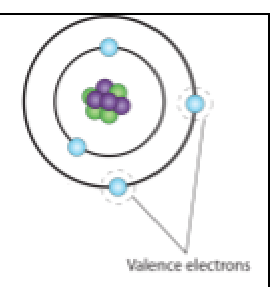
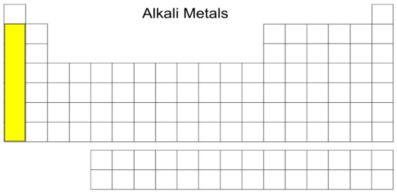
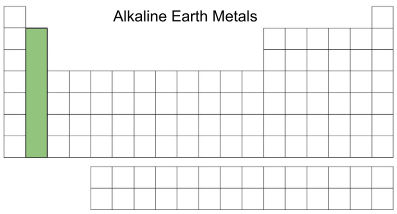



Figure 8

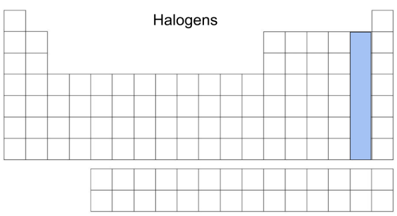
As the elements were organized on the periodic table, a pattern emerged with reactivity as well. Chemical reactivity is based on the outer electrons of an atom, referred to as valence electrons, as shown in figure 8. As the elements move across the period, the number of electrons increased by one. This pattern allowed for elements in a group to have the same number of valence electrons and thus the same reactivity. This holds true for the main group elements (1, 2, 13, 14, 15, 16, 17, 18). The transition metals behave a little differently based on their electron configuration.



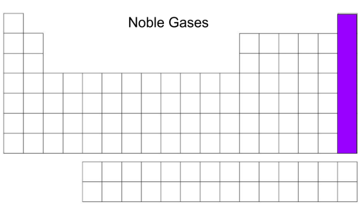

*Group 1: Alkali Metals*
These elements have similar physical properties related to their metallic appearance. They are silver-colored, soft metals that have never been found in elemental form in nature due their high reactivity. Some alkali metals are involved in many biological functions. They have one valence electron that is readily transferred to other atoms.



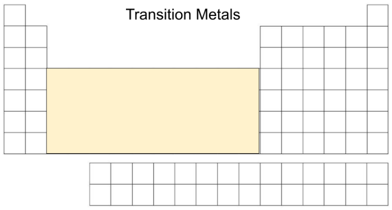

*Group 2: Alkaline Earth Metals*
The metals in this group are also silver-colored, soft metals, but they are slightly less reactive than the alkali metals. Alkaline earth metals are also important in many biological processes. They have two valence electrons that can be readily transferred to other atoms.




*Group 17: Halogens*
The elements found in this group are highly reactive nonmetals that are rarely found in elemental form in nature. They have seven electrons in the outermost shell and will readily gain an electron. Halogens are the only periodic table group that contains elements in all three familiar states of matter at standard temperature and pressure.

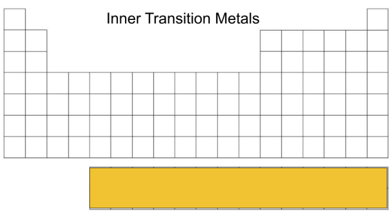



*Group 18: Noble Gases*
The elements in this group consist of nonreactive gases that are always found in elemental form in nature. They are known as the noble gases and are odorless, colorless, and monatomic. They have eight valence electrons in their outer shell, except for helium, which only has two. All the elements in this group have a full energy level, or **octet**, of valence electrons. This full set of valence electrons is what makes these elements nonreactive, as they have very little or no tendency to gain or lose electrons. There are many modern-day applications for noble gases. As these gases emit different colors of light when exposed to electrical discharge, they are often used as decorative lighting in buildings or in signs. Examples include helium, neon, argon, krypton, and xenon.




*Transition Metals*

These elements include the transition metals found in the middle columns of the periodic table, groups 3–12. They have a variable number of outer electrons present in more than one shell, leading to a variety of possible charges for each element. Even though these elements have a variable number of valence electrons, they share many similar chemical and physical properties. As with all metals, the transition elements are both ductile and malleable, are good conductors of both electricity and heat, and have high density and high melting-boiling points. Examples include iron, nickel, copper, gold, and silver. Compounds of transition metals are usually colored.




*Inner Transition Metals*

This includes the lanthanide and actinide series. They are located at the bottom of the periodic table and are called the **inner transition metals**. It is rare to find these elements in a pure form in nature. As a result, they are very expensive to obtain. There are several modern-day industries and applications that depend on rare earth metals, such as the glass-polishing industry, certain computer applications, and production of some magnets. All the actinides are radioactive, and most do not occur naturally on Earth. Instead, they can only be created in laboratories under very specific conditions, and most have incredibly short half-lives.

## Periodic Trends

*Predict Properties with Trends*

As stated earlier, the elements on the periodic table follow periodic law. The characteristics can be used to determine trends. These characteristic properties of elemental families can be attributed to the valence electrons in the outermost electron orbital of each element. All periodic trends can be explained by three simple concepts: attraction of the nucleus for the valence electrons (opposite charges attract), the repulsion of the electrons from each other (like charges repel), and the energy levels within the atom.

*Atomic Radii*

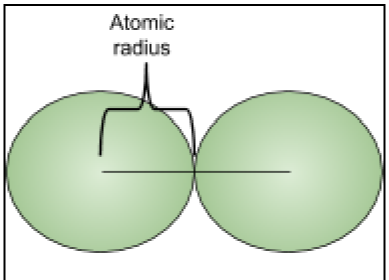


Figure 9

The **atomic radius** is the measurement from the nucleus of one atom to the nucleus of the same atom it is touching and then divided in half. The atomic radii of atoms will decrease when moving from left to right on the periodic table. This is due to the increasing number of protons within the nuclei pulling on the electrons. As you move down a group, the atomic radii will increase. This is due to the addition of electron orbitals. With more electrons farther away from the nucleus, there is greater repulsion, resulting in a larger radius. Figure 10 shows the overall trend in atomic radii on the periodic table.

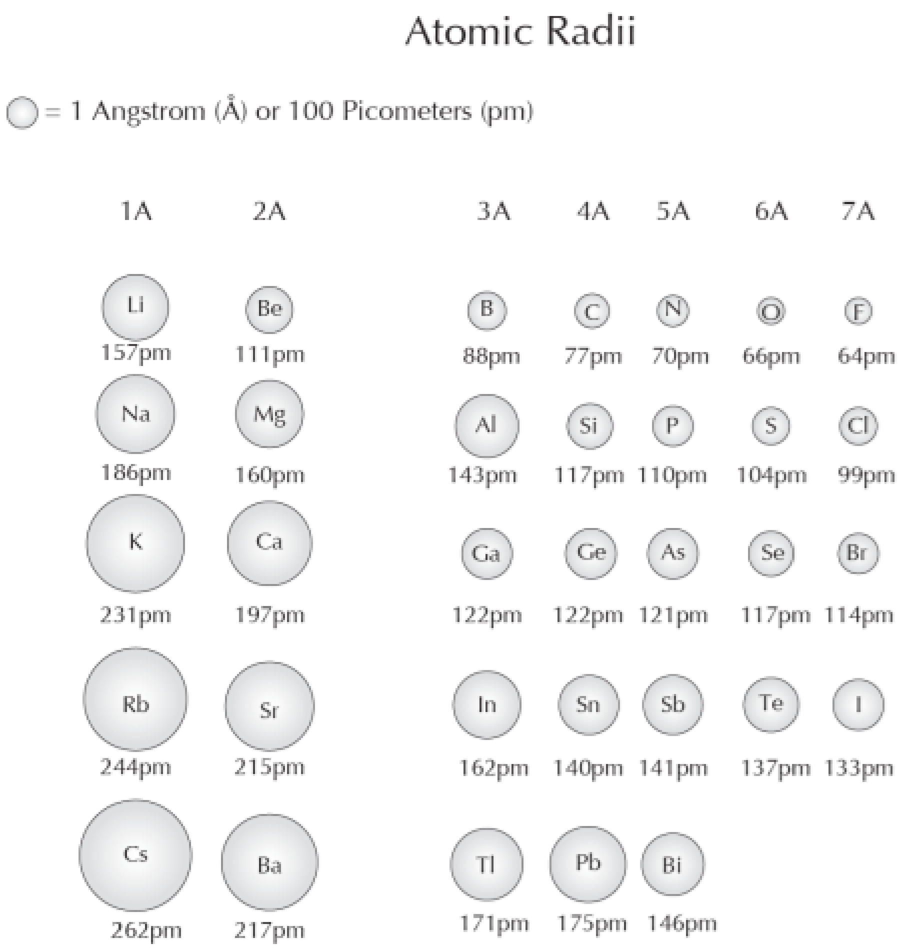



Figure 10

*Ions*

Atoms are neutral because they have the same number of protons and electrons. However, based on valence electrons, some atoms will easily gain or lose electrons to other atoms, creating an **ion**. An ion is an atom with a charge. This charge is determined by the number of electrons either gained or lost. In the case of atoms in group 1, they easily transfer their one valence electron to other elements. This creates an atom with more protons than electrons, thus a positively charged ion called a cation.

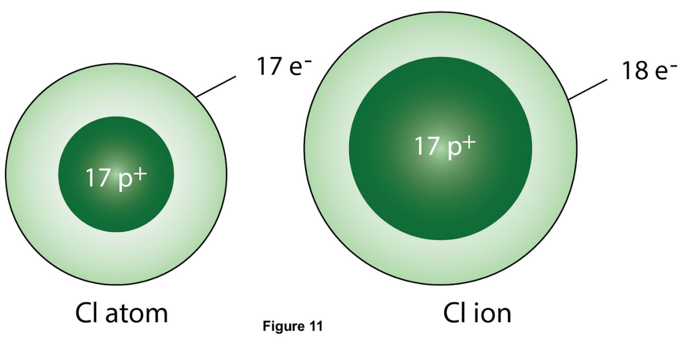



Figure 11

The opposite occurs when an atom gains an electron. In this case, the atom now has more electrons than protons, creating a negatively charged ion called an anion. Figure 11 shows a chlorine atom and a chlorine ion.

*Ionization Energy*

Another important periodic trend is ionization energy, which is the amount of energy required to remove an electron from a neutral atom. The ionization energy of elements within the same chemical family decreases with increasing atomic number. Ionization energy generally increases as you move from left to right across a period and as you move up a group.

*Ionic Radii*

As with the atomic radius, there is a trend with the ionic radius. As you move down a group, in general, the **ionic radii** will increase due to the addition of an electron orbital. As you go across a period, there is a decrease in the ionic radii of cations. Then you will see a jump in the ionic radii as anions are formed. But as you continue across the period, the anion radii will decrease as well. Refer to figure 12.

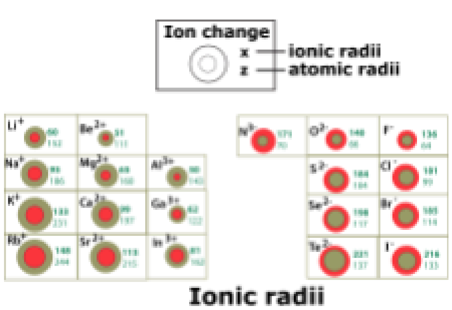



Figure 12

This trend makes sense when you look at the forces occurring in the ions. The positively charged ion will become smaller due to larger attractive forces between the protons in the nucleus and the electrons. The negatively charged ion will become larger due to an increased number of electrons and the resulting electron repulsion, thereby increasing the ionic radius.

**Periodic Table of the Elements**

Decreasing Electronegativity

| | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 1 | 1 **H** 2.20 | | | | | | | | | | | | | | | | | |
| 2 | 3 **Li** 0.98 | 4 **Be** 1.57 | | | | | | | | | | | 5 **B** 2.04 | 6 **C** 2.55 | 7 **N** 3.04 | 8 **O** 3.44 | 9 **F** 3.98 | |
| 3 | 11 **Na** 0.93 | 12 **Mg** 1.31 | | | | | | | | | | | 13 **Al** 1.61 | 14 **Si** 1.90 | 15 **P** 2.19 | 16 **S** 2.58 | 17 **Cl** 3.16 | |
| 4 | 19 **K** 0.82 | 20 **Ca** 1.00 | 21 **Sc** 1.36 | 22 **Ti** 1.54 | 23 **V** 1.63 | 24 **Cr** 1.66 | 25 **Mn** 1.55 | 26 **Fe** 1.83 | 27 **Co** 1.88 | 28 **Ni** 1.91 | 29 **Cu** 1.90 | 30 **Zn** 1.65 | 31 **Ga** 1.81 | 32 **Ge** 2.01 | 33 **As** 2.18 | 34 **Se** 2.55 | 35 **Br** 2.96 | |
| 5 | 37 **Rb** 0.82 | 38 **Sr** 0.95 | 39 **Y** 1.22 | 40 **Zr** 1.33 | 41 **Nb** 1.6 | 42 **Mo** 2.16 | 43 **Tc** 1.9 | 44 **Ru** 2.2 | 45 **Rh** 2.28 | 46 **Pd** 2.20 | 47 **Ag** 1.93 | 48 **Cd** 1.69 | 49 **In** 1.78 | 50 **Sn** 1.96 | 51 **Sb** 2.05 | 52 **Te** 2.1 | 53 **I** 2.66 | |
| 6 | 55 **Cs** 0.79 | 56 **Ba** 0.89 | 57 **La** 1.1 | 72 **Hf** 1.3 | 73 **Ta** 1.5 | 74 **W** 2.36 | 75 **Re** 1.9 | 76 **Os** 2.2 | 77 **Ir** 2.20 | 78 **Pt** 2.28 | 79 **Au** 2.54 | 80 **Hg** 2.00 | 81 **Tl** 1.62 | 82 **Pb** 2.33 | 83 **Bi** 2.02 | 84 **Po** 2.0 | 85 **At** 2.2 | |
| 7 | 87 **Fr** 0.7 | 88 **Ra** 0.9 | 89 **Ac** 1.1 | | | | | | | | | | | | | | | |

Increasing Electronegativity

**Figure 13**

*Electronegativity*

The **electronegativity** of an element reflects its ability to attract electrons to itself in a chemical bond. Electronegativity generally increases when moving from left to right across a period and decreases when moving down a group. This makes fluorine the most electronegative element and francium the least electronegative element.

## Atomic Bonding

When stating that atoms lose or gain electrons, the reality is that electrons are being transferred or shared with other atoms, creating a bond. In atomic bonding, there are three main types of bonds: ionic, covalent, and metallic. The nature of these bonds is based on the behavior of the electrons in each atom.

*Ionic Bonding*

In **ionic bonds,** electrons are generally transferred from a metal to a nonmetal, creating an ion. The atom that loses electrons becomes a positively charged ion, or cation. The atom that gains electrons becomes a negatively charged ion, or anion. The attraction between the newly formed cation and anion results in the formation of an ionic bond. The basic unit of ionic bonds is known as a formula unit. Therefore, the formula unit of sodium chloride is a sodium cation and a chloride anion.

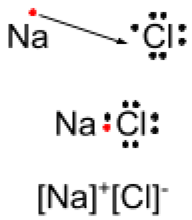



Figure 14

Figure 14 shows the ionic bond formation between sodium and chlorine. Sodium has one valence electron, while chlorine has seven valence electrons. The goal is to reach an octet, so sodium will transfer its one valence electron to chlorine, creating a cation (sodium) and an anion (chlorine). The two ions will attract each other, creating an ionic compound: sodium chloride.

*Covalent Bonding*

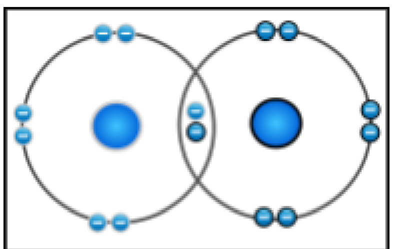


Figure 15

In **covalent bonds**, the valence electrons of each atom are shared and localized between two atoms (figure 15). Starting with the central atom, the electrons are placed around each atom in order to fulfill the octet rule. Important exceptions to the octet rule include hydrogen and helium, as both atoms only require two valence electrons to have a full outer electron orbital.

Sharing two electrons between two atoms is considered a single bond. In many cases, however, covalent compounds cannot form through single bonding.

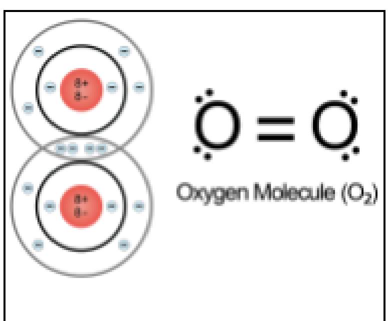



Figure 16

*Double Bonds*

Molecules can also form between atoms that share more than one electron pair. For example, in a molecule of oxygen gas ($O_2$), each oxygen atom (O) has six valence electrons, so each atom needs two electrons to complete its valence orbital (figure 16). Thus, two electron pairs are shared between the two oxygen atoms. The sharing of two electron pairs between atoms creates a double bond that is represented by two parallel lines in Lewis diagrams.

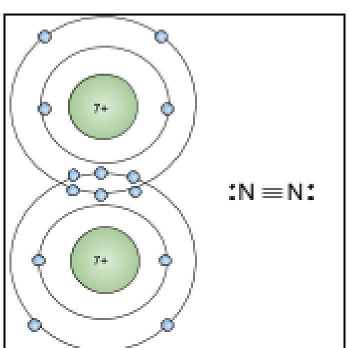



Figure 17

*Triple Bonds*

In another example, the nitrogen atoms (N) in a molecule of nitrogen gas ($N_2$) share three electron pairs (figure 17). This type of covalent bond is known as a triple bond. Double and triple bonds can also occur between different elements. Triple bonds are represented by three parallel lines in Lewis diagrams.

*Trends and Bonding*

Most often, ionic bonds occur between metals and nonmetals, while covalent bonds occur between two nonmetals. The electronegativity of the elements, however, can actually provide a more accurate prediction of the type of bond. To determine what type of bond will occur, you will need to look at the difference in electronegativities. If the electronegativity difference is greater than 1.67, the bond will be ionic, meaning the electrons have been transferred between the elements, creating ions and an ionic bond. If the electronegativity is less than 0.2, the electrons will be shared equally and considered nonpolar covalent. If the difference in electronegativity is between 0.2 and 1.6, the electrons will be shared unequally and considered polar covalent.

*Metallic Bonding*

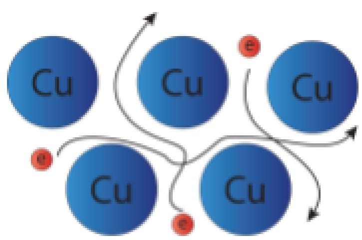


Figure 18

In **metallic bonding**, the electrons in the substance are delocalized, which means that they do not remain attached to any one atom. Metallic bonding is formed from the attraction of the electrons to the metal cations. The valence electrons of metals, while in the solid state, are free to move around the cations within the metals, as shown in figure 18. This freedom of electron movement forms something called the electron sea. Examples of substances that contain metallic bonds include gold bars, sheets of aluminum foil, iron pans, and copper wires.

As a result of the electrons' ability to move freely among the cations, metals are good thermal and electrical conductors. The structure of the electron sea in a metallic bond provides many other properties observed in metals. The abilities of the electrons to move throughout the electron sea and of the cations to slide past each other make metallic bonds more flexible than ionic or covalent bonds. This flexibility makes metals both malleable (able to be hammered and bent into shapes) and ductile (able to be stretched into long, thin wires).

The properties of elements are determined by their structure. This structure allows for organization on the periodic table. The periodicity of the elements helps predict how elements will interact with each other, including the number and types of bonds.

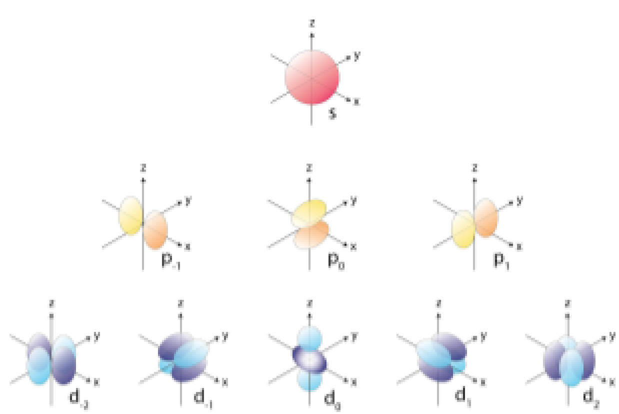
# Advanced Topics

*Electron Configuration*

Further study provided more information to help predict the location of electrons based on energy. Ultimately, the electrons were organized based on their proximity to the nucleus and the shape of the orbital in which they traveled. This organization included a principal quantum number, a sublevel, the number of orbitals within that sublevel, and the spin within that orbital. With this organization, scientists could account for where an electron would most likely be found in the electron cloud.

The principal quantum number, indicated by *n*, relates to the main energy level an electron could possibly occupy. This quantum number currently starts at 1 and goes up to 7. As *n* increases, so does the energy of the electron as well as the distance from the nucleus. So, if an electron has a principal quantum number of 1, it would occupy the first energy level, with the lowest amount of energy and closest to the nucleus. An electron with a principal quantum number of 7 would have the higher amount of energy and be farthest from the nucleus.

Within each energy level, there are sublevels that represent different-shaped orbitals. These have a designation of *s*, *p*, *d*, or *f*. As *n* increases, the number of sublevels will increase. The figure below represents the different shapes that *s*, *p*, and *d* orbitals are described as. The *f* orbital is very complicated and hard to represent graphically. The shapes in general represent the likelihood of where electrons are found within that sublevel.



Orbital shapes within sublevels

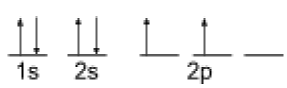
Refer to the table below to see the overall organization of quantum numbers used for electron configuration.

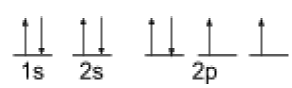
| Principal Quantum Number (Energy Level) | Sublevels | Number of Orbitals | Number of Electrons per Sublevel | Total Number of Electrons per Energy Level |
|---|---|---|---|---|
| 1 | s | 1 | 2 | 2 |
| 2 | s | 1 | 2 | 8 |
| | p | 3 | 6 | |
| 3 | s | 1 | 2 | 18 |
| | p | 3 | 6 | |
| | d | 5 | 10 | |
| 4 | s | 1 | 2 | 32 |
| | p | 3 | 6 | |
| | d | 5 | 10 | |
| | f | 7 | 14 | |

When writing electron configurations, there are guidelines that must be remembered.

1. Electrons will always fill the lowest energy level first.
2. The Pauli exclusion principle states that an orbital can only contain, at most, two electrons, and those electrons must have an opposite spin.
3. Hund's rules state that electrons must fill orbitals of the same energy one at time before a second electron will enter.

Using these guidelines, here are two ways to represent the electron configurations for a few elements:

Carbon $1s^2\ 2s^2\ 2p^2$ 

Oxygen $1s^2\ 2s^2\ 2p^4$ 


There are two electrons in the 1s and 2s. This is represented by two arrows pointing in a different direction. This does not mean that the electrons point up and down, but rather it means that their spins are opposite each other (Pauli exclusion principle). There are two electrons in 2p; based on Hund's rules, an electron is placed in each orbital before the second electron will fill that orbital. This is represented by placing one arrow in each orbital. In contrast, you can see that oxygen's representation has four electrons in 2p. One arrow was placed in each orbital before the second down-spin arrow was added to the first 2p orbital.

# Beyond the Classroom

Dmitri Mendeleev had a huge task trying to arrange the 25 elements known at that time into a usable manner that made sense and could be understood by all those who would need that information. The information he had available included just the physical and chemical properties of each element.

You are now going to put yourself in Mendeleev's shoes. Instead of elements, you will design your periodic table of foods.

1. List your 20 favorite foods. Be specific.
2. Organize these foods into groups and periods.
3. Provide a name for each group.
4. Be able to explain how your foods were organized so others could use your periodic table of food, if needed.
5. Is there any other format you can have used to organize your table, such as a circular chart? If so, explain what other format you would use and why it would work better than a traditional table with rows and columns.

# Periodic Table and Element Structure Review

**Reviewing Key Terms**

Use each of the following terms in a separate sentence.

1. Valence electrons
2. Ionic bonding
3. Nucleus
4. Period

Use the correct key term to complete each of the following sentences.

1. _______ is the name for the columns on the periodic table.
2. _______ are characterized as being brittle.
3. Most of the elements on the periodic table are classified as __________.

**Reviewing Main Ideas**

1. Which element has the smallest radius?
   a. Fluorine
   b. Iodine
   c. Chlorine
   d. Bromine

2. Which of the following represents the subparticles for magnesium?
   a. 12p, 12e, 13n
   b. 12p, 13e, 14n
   c. 12p, 13e, 14n
   d. 12p, 12e, 12n

3. Which type of bond will share valence electrons between two atoms?
   a. Ionic
   b. Covalent
   c. Metallic
   d. All of these

**Making Connections**

1. Explain the difference between an ionic compound and a covalent compound, and give two elements that would combine to form each.
2. Describe the difference between $Al^{3+}$ and $N^{3-}$.

**Open-Ended Response**

1. Explain how metallic bonding supports the properties of metals.
2. Describe at least two types of tests you could perform to determine if a substance was a metal or nonmetal.
3. Choose a group from the periodic table that you encounter the most in your everyday life.
4. Compare and contrast $^{1}H$, $^{2}H$, and $^{3}H$.

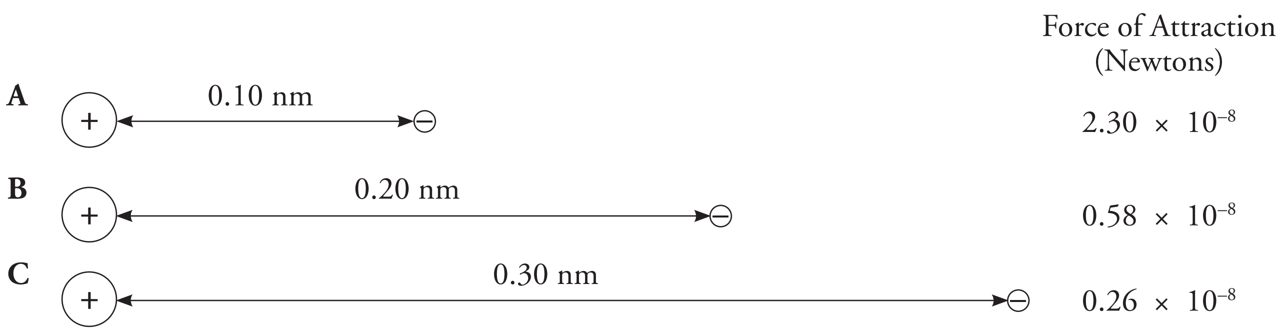
# Coulombic Attraction

What variables will affect the force of attraction between charged particles?

## Why?

**Coulombic attraction** is the attraction between oppositely charged particles. For example, the protons in the nucleus of an atom have attraction for the electrons surrounding the nucleus. This is because the protons are positive and the electrons are negative. The attractive force can be weak or strong. In this activity, you will explore the strength of attraction between protons and electrons in various atomic structures.

## Model 1 – Distance and Attractive Force




1. What subatomic particles do these symbols represent in Model 1?

   (+)  ⊖

2. Would you expect to observe attraction or repulsion between the subatomic particles in Model 1?

3. Consider the data in Model 1.

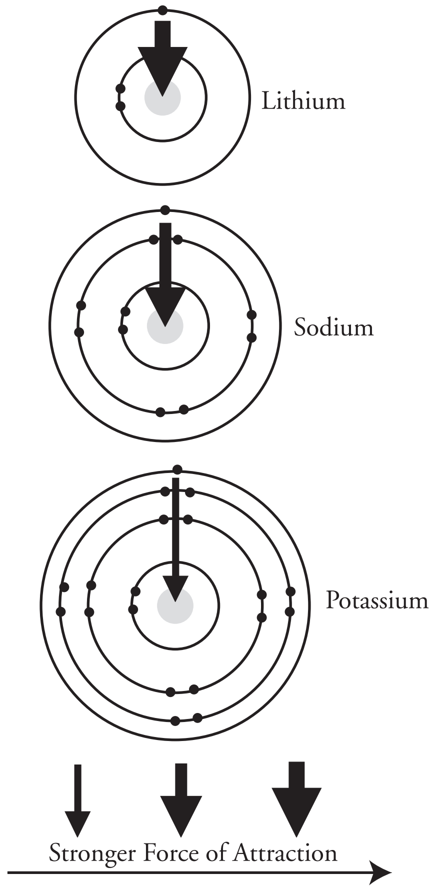
   *a.* What are the independent and dependent variables in the data?

   *b.* Write a complete sentence that describes the observed relationship between the independent and dependent variables in Model 1.

4. If the distance between a proton and electron is 0.50 nm, would you expect the force of attraction to be greater than or less than $0.26 \times 10^{-8}$ N?

5. If two protons are 0.10 nm away from one electron, would you expect the force of attraction to be greater than or less than $2.30 \times 10^{-8}$ N?

## Model 2 – The Alkali Metals




6. Consider the diagrams in Model 2.
   *a.* What do the arrows represent?

   *b.* How does the thickness of the arrows relate to the property given in part *a*?

7. Using a periodic table, locate the elements whose atoms are diagrammed in Model 2. Are the elements in the same column or the same row?

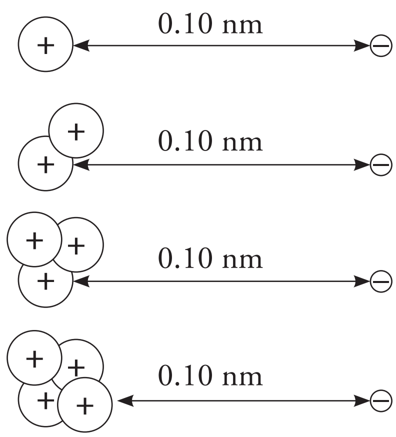
8. Circle the outermost electron in each of the diagrams in Model 2.
   *a.* As you move from the smallest atom to the largest atom in Model 2, how does the distance between the outermost electron and the nucleus change?

   *b.* As you move from the smallest atom to the largest atom in Model 2, how does the attractive force between the outermost electron and the nucleus change?

   *c.* Are your answers to parts *a* and *b* consistent with the information in Model 1?

STOP

## Model 3 – Number of Protons and Attractive Force




| | Force of Attraction (Newtons) |
|---|---|
| **A** | $2.30 \times 10^{-8}$ |
| **D** | $4.60 \times 10^{-8}$ |
| **E** | $6.90 \times 10^{-8}$ |
| **F** | $9.20 \times 10^{-8}$ |

9. Consider the data in Model 3.
   *a.* What are the independent and dependent variables in the data?

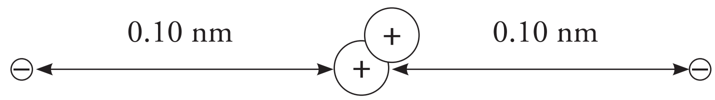
   *b.* Write a complete sentence that describes the relationship between the independent and dependent variables in Model 3.

10. What would be the attractive force on a single electron if five protons were in the nucleus of an atom? Show mathematical work to support your answer.

11. Imagine that a second electron were placed to the left of a nucleus containing two protons (Model 3, set D). Predict the force of attraction on both the original electron and the second electron. Explain your prediction with a complete sentence.

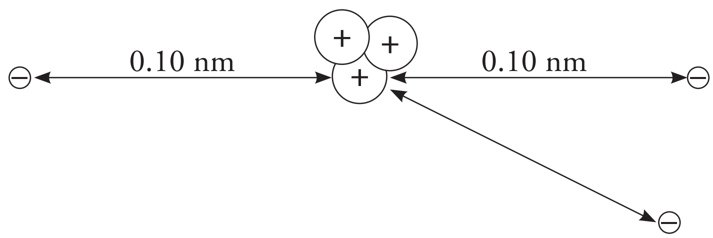
## Read This!

The attractive and repulsive forces in an atom are rather complex. An electron is attracted to the protons in the nucleus, but it is also repelled by the other electrons in the atom. It is important to note however that the attractive force of the nucleus is NOT divided up among the electrons in the atom. Each electron gets approximately the full attractive force of the nucleus (minus the repulsive effects of other electrons). Compare the diagram below to set D in Model 3. Notice the similarity in attractive force.

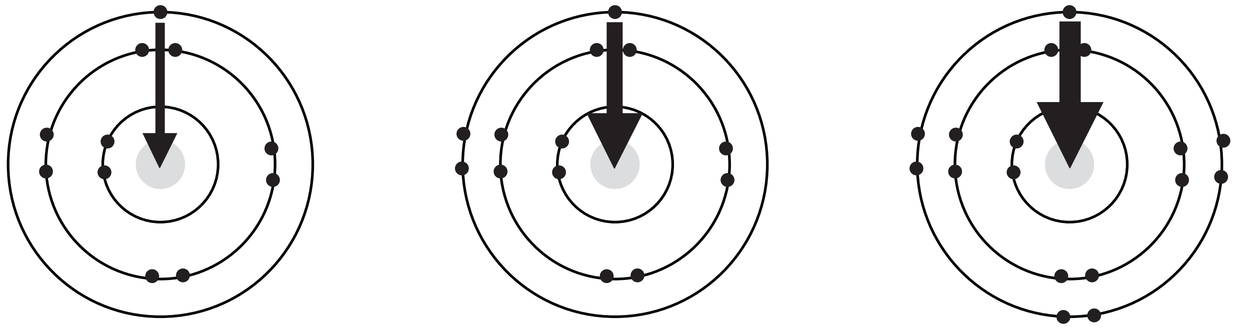



approx. $4.60 \times 10^{-8}$
(on each electron)

12. What is the approximate attractive force on each electron below?



## Model 4 – Period 3 Elements



Sodium Aluminum Chlorine

13. Using the periodic table, locate the elements whose atoms are diagrammed in Model 4. Are the elements in the same column or the same row?

14. Circle the outermost electron(s) in each of the atoms in Model 4.

15. Which of the three atoms diagrammed in Model 4 has the strongest attraction for its outermost electron(s)?

16. Consider the information in Model 4.

    *a.* As you move from the smallest atom to the largest atom, does the distance between the outermost electron(s) and the nucleus change significantly?

    *b.* Can the differences in the attractive force shown by the arrows be explained by a change in the distance between the electron(s) and the nucleus?

    *c.* On the diagrams in Model 4, write the number of protons located in the nucleus of each atom.

    *d.* Can the differences in attractive forces shown by the arrows in Model 4 be explained by a change in the number of protons in the nucleus? If yes, explain the relationship in Model 4.

STOP

17. For each set of elements below, circle the element whose atoms will have a stronger attractive force between their outermost electron(s) and the nucleus.

    a. Ba and Ca *b.* Cr and Cu *c.* Ar and Xe

## Extension Questions

18. Consider the atom diagrams in Model 2.

    *a.* On each diagram write the number of protons in the nucleus of the atom.

    *b.* When comparing elements in the same column of the periodic table, which factor—distance to the nucleus or number of protons in the nucleus—seems to be the dominant factor for determining the attractive force between the outermost electron(s) and the nucleus? Explain.

19. Consider the data presented in Models 1 and 3.

    *a.* Describe the mathematical relationship between the distance ($d$) and the attractive force ($F$) between protons and electrons.

    *b.* Describe the mathematical relationship between the number of protons in the nucleus ($Z$) and the attractive force ($F$) between the nucleus and electrons.

# Periodic Trends

Can the properties of an element be predicted using a periodic table?

## Why?

The periodic table is often considered to be the "best friend" of chemists and chemistry students alike. It includes information about atomic masses and element symbols, but it can also be used to make predictions about atomic size, electronegativity, ionization energies, bonding, solubility, and reactivity. In this activity you will look at a few periodic trends that can help you make those predictions. Like most trends, they are not perfect, but useful just the same.

1. Consider the data in Model 1 on the following page.
   *a.* Each element has three numbers listed under it. Which value represents the atomic radius?

   *b.* What are the units for the atomic radius?

   *c.* Write a complete sentence to convey your understanding of atomic radius. *Note:* You many not use the word "radius" in your definition.

2. In general, what is the trend in atomic radius as you go down a group in Model 1? Support your answer, using examples from three groups.

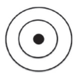
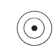
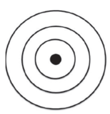
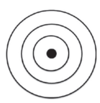
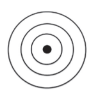
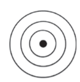
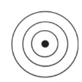
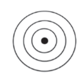
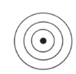
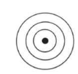
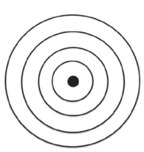
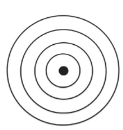
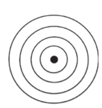
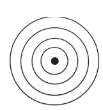
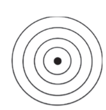
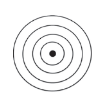
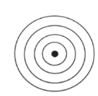
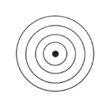
3. Using your knowledge of Coulombic attraction and the structure of the atom, explain the trend in atomic radius that you identified in Question 2. *Hint:* You should discuss either a change in distance between the nucleus and outer shell of electrons or a change in the number of protons in the nucleus.

4. In general, what is the trend in atomic radius as you go across a period (left to right) in Model 1? Support your answer, using examples from two periods.

5. Using your knowledge of Coulombic attraction and the structure of the atom, explain the trend in atomic radius that you identified in Question 4.

STOP

## Model 1 – Main Group Elements

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| 1 | | | | | | | 2 |
| H | | | | | | | He |
|  | | | | | | |  |
| 37 | | | | | | | 31 |
| 1312 | | | | | | | 2372 |
| 2.1 | | | | | | | N/A |
| 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 |
| Li | Be | B | C | N | O | F | Ne |
|  |  |  |  |  |  |  |  |
| 152 | 112 | 83 | 77 | 71 | 66 | 71 | 70 |
| 520 | 900 | 801 | 1086 | 1402 | 1314 | 1681 | 2081 |
| 1.0 | 1.5 | 2.0 | 2.5 | 3.0 | 3.5 | 4.0 | N/A |
| 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
| Na | Mg | Al | Si | P | S | Cl | Ar |
|  |  |  |  |  |  |  |  |
| 186 | 160 | 143 | 117 | 115 | 104 | 99 | 98 |
| 496 | 738 | 578 | 786 | 1011 | 1000 | 1251 | 1521 |
| 0.9 | 1.2 | 1.5 | 1.8 | 2.1 | 2.5 | 3.0 | N/A |
| 19 | 20 | 31 | 32 | 33 | 34 | 35 | 36 |
| K | Ca | Ga | Ge | As | Se | Br | Kr |
|  |  |  |  |  |  |  |  |
| 227 | 197 | 122 | 123 | 125 | 117 | 114 | 112 |
| 404 | 550 | 558 | 709 | 834 | 869 | 1008 | 1170 |
| 0.8 | 1.0 | 1.7 | 1.8 | 1.9 | 2.1 | 2.5 | N/A |

| |
|---|
| Atomic Number |
| Element Symbol |
| Electron Shell Diagram |
| Atomic Radius (pm) |
| 1st Ionization Energy (kJ/mol) |
| Electronegativity |

*Note:* The transition elements and f-block elements have been removed from the periodic table here to ease the analysis of the trends.

6. Locate the numbers in Model 1 that represent the ionization energy. The **ionization energy** is the amount of energy needed to remove an electron from an atom.
   *a.* Using your knowledge of Coulombic attraction, explain why ionization—removing an electron from an atom—takes energy.

   *b.* Which takes more energy, removing an electron from an atom where the nucleus has a tight hold on its electrons, or a weak hold on its electrons? Explain.

7. In general, what is the trend in ionization energy as you go down a group? Support your answer using examples from three groups.

8. Using your knowledge of Coulombic attraction and the structure of the atom, explain the trend in ionization energy that you identified in Question 7.

9. In general, what is the trend in ionization energy as you go across a period? Support your answer using examples from two periods.

10. Using your knowledge of Coulombic attraction and the structure of the atom, explain the trend in ionization energy that you identified in Question 9.

11. Atoms with loosely held electrons are usually classified as metals. They will exhibit high conductivity, ductility, and malleability because of their atomic structure. Would you expect metals to have high ionization energies or low ionization energies? Explain your answer in one to two complete sentences.

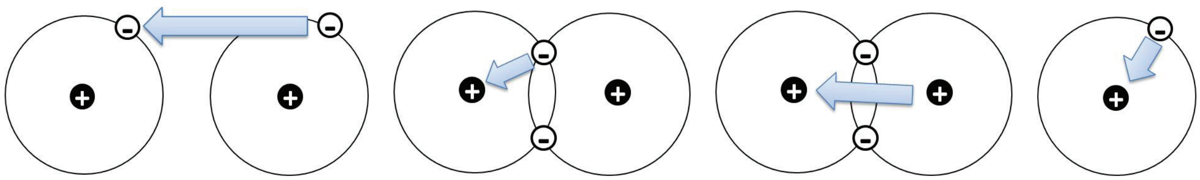
STOP

## Read This!

**Electronegativity** is a measure of the ability of an atom's nucleus to attract electrons from a different atom within a covalent bond. A higher electronegativity value correlates to a stronger pull on the electrons in a bond. This value is only theoretical. It cannot be directly measured in the lab.

12. Using the definition stated in the *Read This!* box above, select the best visual representation for electronegativity. Explain your reasoning.

A B C D



13. Locate the electronegativity values in Model 1.

    *a.* What is the trend in electronegativity going down a group in Model 1?

    *b.* Explain the existence of the trend described in part *a* in terms of atomic structure and Coulombic attraction.

    *c.* What is the trend in electronegativity going across a period in Model 1?

    *d.* Explain the existence of the trend described in part *c* in terms of atomic structure and Coulombic attraction.

14. The two diagrams below can summarize each of the three trends discussed in this activity. Write "atomic radius," "ionization energy," and "electronegativity" under the appropriate diagram.

## Extension Questions

15. During this activity you may have noticed that not all of the data provided in the models followed the trends.

    *a.* Identify two places in Model 1 where the property listed does not fit the trend identified in this activity.

    *b.* Why is it still beneficial for chemists to understand as many periodic trends as they can?

    *c.* Propose an explanation for one of the exceptions you identified in part *a*. Use your knowledge of atomic structure and Coulombic attraction in your hypothesis.

16. Rank the following elements from **smallest to largest** electronegativity based on the trends you have discovered thus far in the periodic table: barium (atomic number 56), bromine (atomic number 35), and iron (atomic number 26). Explain your reasoning.