When you hear the word "Chemistry", *what comes to mind?* There may be a lot of things, but one of those is probably the periodic table of the elements! In this chapter, we will talk about how elements are arranged in the modern periodic table, and what important properties and trends can we obtain from this arrangement.

The Periodic Table of Elements

The development of the periodic table as we know it today started in 1829 when **Johann Dobereiner** proposed the model of **triads** or the law of triads in some references. The model states that the atomic weight of the middle element in a set of three elements (he called triad) is approximately equal to the arithmetic mean of the two extreme elements in the set.

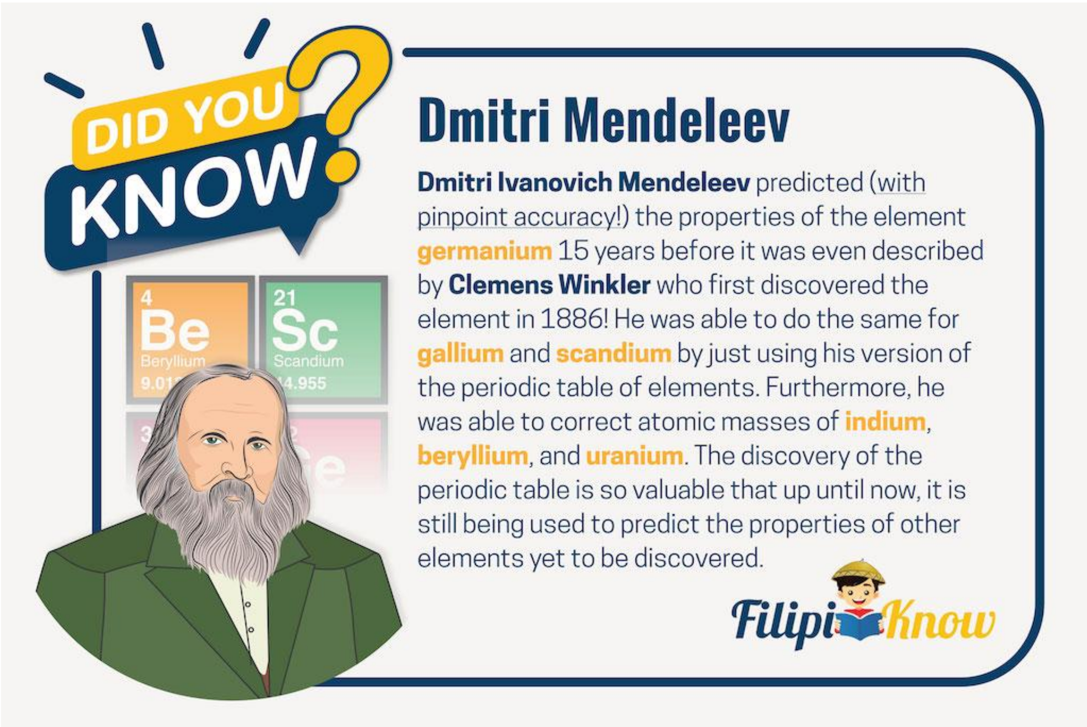
Dobereiner was able to identify several triads of elements with the same properties which adhere to this model. Unfortunately, it fails to apply to most of the elements known during Dobereiner's time, proving the model severely lacking. Table 1 below enumerates some of Dobereiner's triad.

Element	Symbol	Atomic Mass
Lithium	Li	6.9
Sodium	Na	23.0
Potassium	K	39.0
Chlorine	Cl	35.4
Bromine	Br	79.9
Iodine	I	126.9
Sulfur	S	32.0

Selenium	Se	79.0
Tellurium	Te	128.0

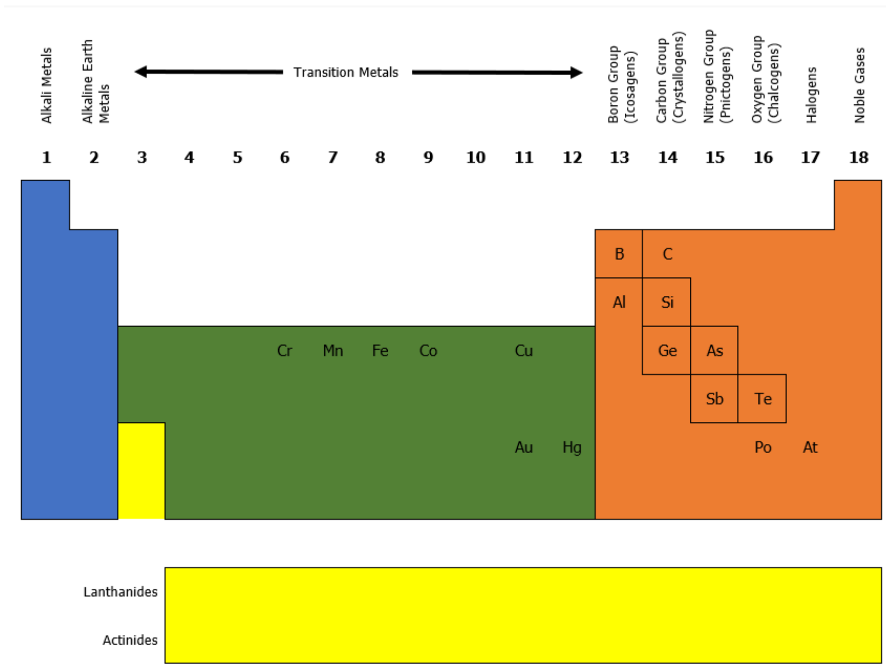
Dobereiner's model was followed by **John Newland**'s **law of octaves** in 1864. Newland observed that when elements are arranged based on their atomic weights, elemental properties seem to repeat every eight elements. Although the law was able to successfully group a few elements with similar properties, it was generally considered unsuccessful.

It was only during the time of **Julius Lothar Meyer** and **Dmitri Ivanovich Mendeleev** that the idea of a periodic table of elements was presented to the world. The two scientists independently formulated the table, but Mendeleev was given most of the credit because of his insistence on how important the table is.



Groups and Families in the Periodic Table

Unlike Mendeleev's table which is arranged based on atomic mass, the modern periodic table that we know today is arranged based on atomic number. The rows in the table are known as periods, while the columns are known as groups or families. The table below shows the outline of the modern periodic table of elements.

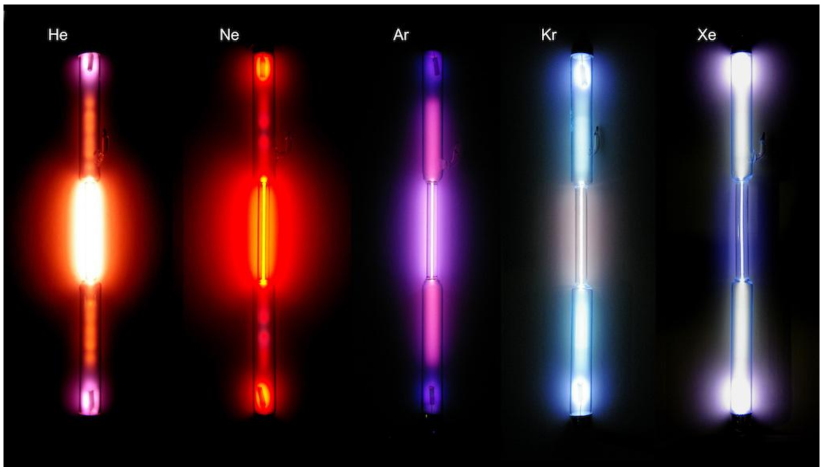


Elements in groups 1, 2, and 3-17 are called the **representative elements** or the **main group elements**. All the representative elements have their highest *s* or *p* orbital incompletely filled.

Belonging to the representative elements, group 1 is known as the **alkali metals**. Elements under this family are highly reactive and possess a single electron in their outermost shell. As a result, these elements readily lose one electron to form ions with a +1 charge. Group 2, on the other hand, is known as **alkaline earth metals**. Elements under this family have two electrons in their valence shell which they readily lose to form cations with +2 charges.

Together with the alkali metals, groups 1 and 2 (except H and Be) form strong inorganic bases when bonded with hydroxide ions. Furthermore, these two groups are sometimes referred to as the ***s block*** elements.

Group 18 (**noble gases**) have completely filled *p* subshells (except He). Aside from helium, all members of the group have 8 electrons in their valence shell. As a result, noble gases are chemically inert. They are also colorless, although ionizing them may cause these gases to emit bright colors. This is the reason why noble gases (except Rn) are commonly used in neon signs.



Groups 3-12 comprise the **transition metals**. Elements belonging to these groups are generally characterized by partially filled *d* subshells or those which can give rise to cations with an incomplete *d* subshell. Because of this, transition metals are sometimes called the ***d* block elements**.

Some transition metals can form cations with multiple oxidation states. For example, chromium can have an oxidation state of either +3 or +6, while manganese can take +2, +3, +4, +6, and +7 as its oxidation states. Other transition metals capable of having multiple oxidation states are labeled in the modern periodic table of elements shown previously.

Group 13 is known as the **Boron group** (or icosagens), which all have three electrons in their outermost shell. This group forms cations with a +3 charge upon ionization.

Carbon group or crystallogens are those under group 14. Elements under this group are characterized by 4 electrons in their outermost shell.

Group 15 are the **pnictogens**, all of which have 5 valence electrons. Aside from nitrogen, pnictogens are solid at room temperature.

Chalcogens or the group 16 elements are composed of electronegative nonmetals and metalloids, all of which have 6 electrons in their outermost shell. Next to chalcogens are the **halogens** (Group 17), famously known as the only group in the periodic table with members existing as gas (fluorine and chlorine), liquid (bromine), and solid (iodine and astatine) at room temperature. Halogens have 7 electrons in their outermost shell; thus, these atoms readily accept an electron to form anions with a –1 charge.

Groups 13-18 are collectively known as the ***p block* elements**. The *p* block is made up of metals, nonmetals, and the so-called metalloids, or those elements with properties that are intermediate to both metals and nonmetals. B, Si, Ge, As, Sb, Te (enclosed in a square border in the outline of the modern periodic table of elements displayed earlier) are widely regarded as metalloids; Po and At are irregularly accepted; while Al and C are rarely recognized as metalloids. Because of their unique properties, metalloids are used in the manufacturing of semiconductors.

Aside from these families, there are also lanthanide and actinide series in the periodic table, which are collectively called the ***f block* transition metals** due to their incompletely filled *f* subshells.

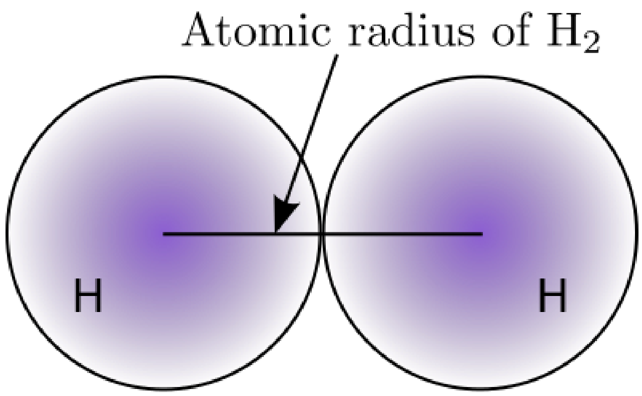
Lanthanides are also known as **rare earth elements**, as they are naturally occurring elements but exist in extremely small amounts (less than 0.001 % of the Earth's crust). Estimates show that lanthanides are actually more abundant than other elements; however, they are highly dispersed in the Earth's crust, which makes these elements sort of "rare". These elements have incompletely filled *4f* subshells or those that form cations with incompletely filled *4f* subshells.

On the other hand, **actinides** are all radioactive. They are heavy elements that are not naturally occurring, some are even named after famous scientists. Elements in the actinide series either have incompletely filled *5f* subshells or form cations with incompletely filled *5f* subshells.

Periodic Trends

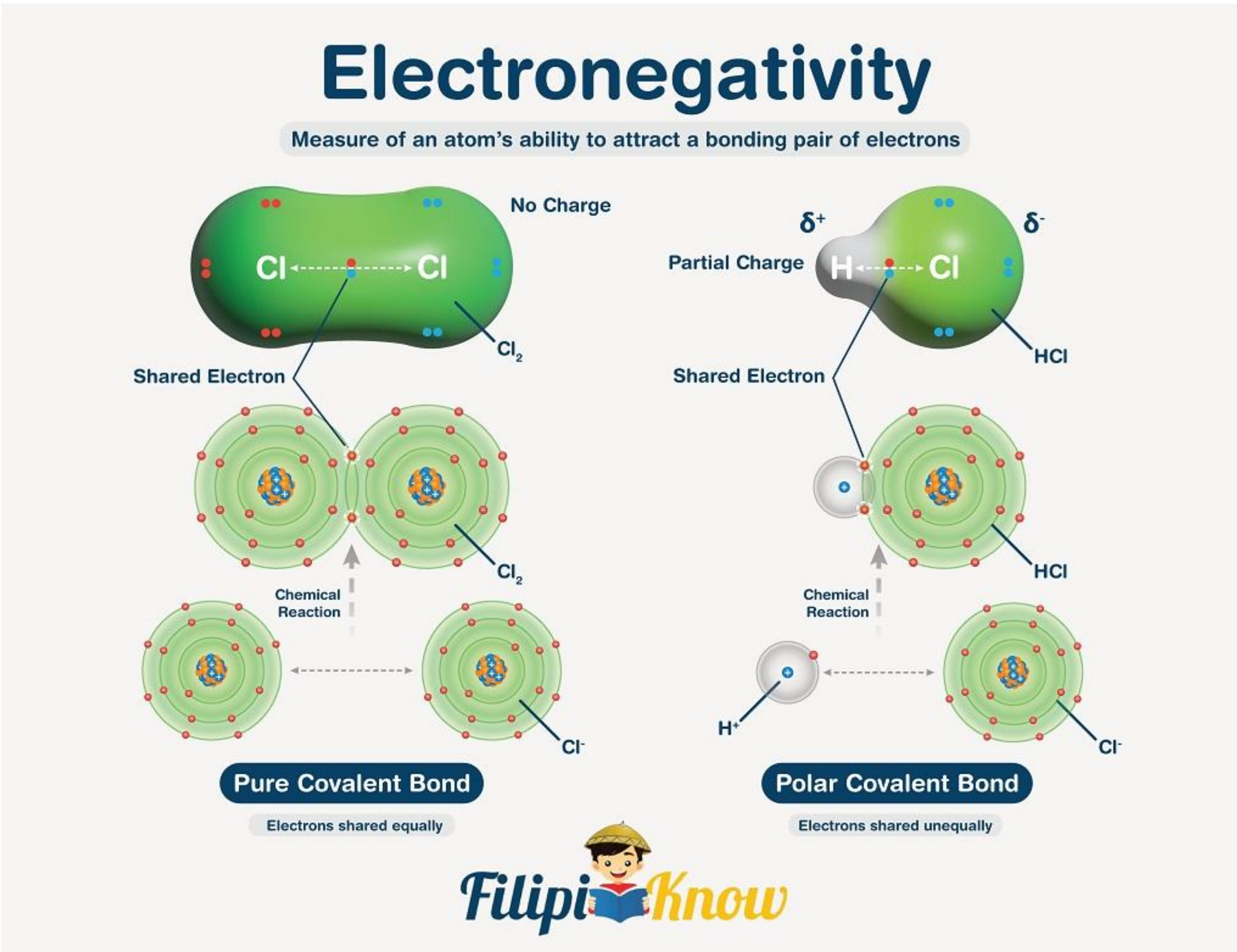
Another benefit of using the periodic table of elements is the ease of determining trends in some of the properties of the elements relative to one another. In this topic, we will discuss trends in atomic radius, electronegativity, ionization energy, and electron affinity.

1. Atomic radius



Atomic radius is defined as one-half the distance between two adjacent atoms. Generally, atomic radius increases down the group and decreases across a period from left to right. This trend can be explained by the fact that as we go down the group, another energy level is being added to the atom. Across a period, elements have the same number of energy levels. However, as we go from left to right, the number of electrons in the outermost shell increases. This increase in the number of electrons causes stronger interaction between the nucleus and the electrons in the valence shell. As a result, the atom becomes “compressed”, and the radius decreases.

2. Electronegativity



Electronegativity is the tendency of an atom to draw shared electrons in a chemical bond towards itself. The figure above shows the difference in electron density between bonded atoms in Cl_2 (left) and HCl (right).

Because Cl atoms have the same electronegativity, you can see that the electron cloud in the two Cl atoms is uniformly distributed. However, in HCl, the difference in electronegativity between H and Cl causes electrons in the covalent bond to be drawn nearer to the Cl atom. As a result, the electron cloud is larger in the vicinity of chlorine compared to the vicinity of hydrogen.

Determining periodic trends in electronegativity is quite easy. You just need to remember that the most electronegative element in the periodic table is fluorine. Hence, electronegativity increases as you approach fluorine (i.e., electronegativity increases across a period from left to right, and decreases down the group).

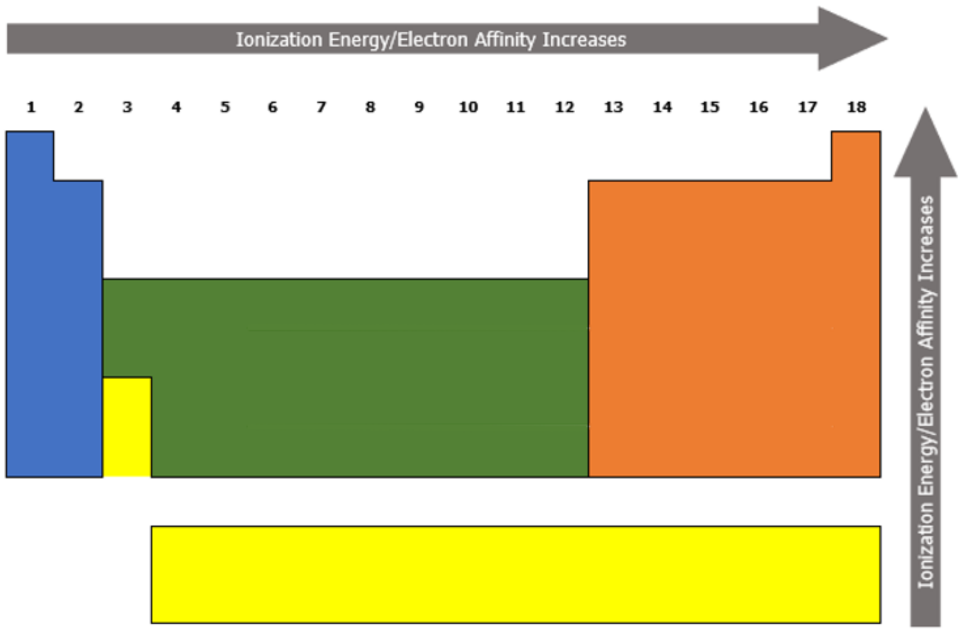
3. Ionization energy

Meanwhile, **ionization energy** is the minimum energy required (in kJ/mol) to remove one electron from an isolated atom or molecule.

One important thing to remember here is that the closer the electron to the nucleus, the stronger the attraction. Hence, electrons closer to the nucleus are harder to remove, or in other words, have higher ionization energy. This follows that removing a valence electron is relatively easier (since they are the one farthest to the nucleus) than removing one of the core electrons in an atom. Because of this, smaller atoms have higher ionization energy, and larger atoms have smaller ionization energy.

Therefore, ionization energy decreases down a group and increases across a period from left to right.

4. Electron affinity



On the other hand, **electron affinity** is the atom's ability to accept one or more electrons. The more willing an atom to accept an electron, the more positive is its electron affinity.

In the periodic table, the halogens or the group 17 elements have the most positive electron affinity. Therefore, electron affinity increases from left to right of the periodic table and decreases down the group.

This trend can also be related to the atomic radius of an atom. The electrons in the valence shell of smaller atoms are closer to the nucleus. As a result, valence electrons are held more tightly by the nucleus in smaller atoms. Upon entering, an incoming extra electron is stabilized by this stronger attraction with the nucleus. In contrast, an incoming electron into the valence shell of larger atoms is not stabilized due to weaker interaction between the nucleus and the electron. As a result, larger atoms have a lower affinity for electrons.