

## Periodic Properties of the Elements

Elements in the same column contain the same number of electrons in their outer-shell orbitals.

Consider oxygen and sulfur:



The similarity in the occupancies of their valences **s** and **p** orbitals leads similarities in their properties.

However, oxygen's outermost electrons are in the 2<sup>nd</sup> shell (n=2) where as sulfur's outermost electrons are in 3<sup>rd</sup> shell (n=3).

The periodic table is the most significant tool that chemist use for organizing chemical facts.

The **periodic law** states that the properties of the elements recur in a repeating pattern when arranged according to increasing atomic number

The number of known elements more than double in the early nineteenth century.

As scientists began to investigate the possibilities of classifying them in useful ways. In 1869 Dmitri Mendeleev in Russia and Lothar Meyer in Germany published nearly identical classification schemes.

Both scientists noted that similar chemical and physical properties recur periodically when the elements are arranged in order of increasing atomic weight.

Mendeleev and Meyer came to essentially the same conclusion about the periodicity of the properties of the elements. However, Dmitri is given credit for advancing his ideas more vigorously and stimulating much new work in chemistry.

His insistence that elements with similar characteristics be listed in the same families forced him to leave several blank spaces in his table.

Mendeleev boldly predicted their existence and properties.

Table 7.1 Comparison of the Properties of Eka-silicon Predicted by Mendeleev with the Observed Properties of Germanium

Property	Mendeleev's Predictions for Eka-silicon (made in 1871)	Observed Properties of Germanium (discovered in 1886)
Atomic weight	72	72.59
Density (g/cm <sup>3</sup> )	5.5	5.35
Specific heat (J/g-K)	0.305	0.309
Melting point (°C)	High	947
Color	Dark gray	Grayish white
Formula of oxide	XO <sub>2</sub>	GeO <sub>2</sub>
Density of oxide (g/cm <sup>3</sup> )	4.7	4.70
Formula of chloride	XCl <sub>4</sub>	GeCl <sub>4</sub>
Boiling point of chloride (°C)	A little under 100	84

John Newlands described an early periodic table in 1864 but was ridiculed because he likened periodicity to the musical octave, even calling it the "Law of Octaves." After Mendeleev's table appeared, Newlands' contributions were finally acknowledged.

In 1913, an English physicist named Henry Moseley developed the concept of atomic numbers.

Moseley determined the frequencies of X-rays emitted after different elements were bombarded with high-energy electrons. He found that each element produces X-rays of a unique frequency.

He found that the frequency generally increased as the atomic mass increased.

He arranged the X-ray frequencies in order by assigning whole number, called an *atomic number*, to each element.

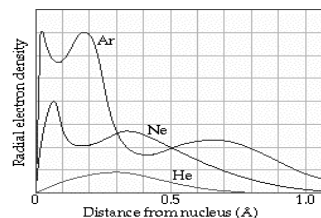
Moseley discovered that the nuclear charge increased by 1 for each element in the periodic table.

Moseley concluded that arranging elements according to increased nuclear charge, rather than atomic mass, more clearly explained the repeating properties of the elements.

### ELECTRON SHELLS AND THE SIZES OF ATOMS

Even before Bohr had proposed his theory of the hydrogen atom, the American chemist Gilbert N. Lewis (1875-1946) had suggested that electrons in atoms are arranged in spherical shells around the nucleus.

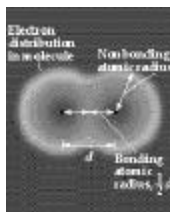
The idea of electron "shells" is useful in visualizing the atom. However, it is a poor descriptor of atomic structure.



The maxima in radial electron density corresponds to Lewis's idea of shells of electrons.

As the nuclear charge increases, the 1s electrons are "pulled" closer and closer to the nucleus.

## Atomic Sizes



The radial-electron-density distributions previously shown do not end abruptly at some distance from the nucleus. Nevertheless, scientists have used a number of means to estimate the radius of an atom, which is called the atomic radius. One of the most common methods to determine atomic radii is to assume that atoms are spheres that touch each other when they are bonded together.

Many properties of molecules depend on the distances between the atoms in the molecule. Atomic radii allow one to estimate the bond lengths between different elements in molecules.

The C-C bond length is 1.54 Å, implying a radius of 0.77 Å for a carbon atom.

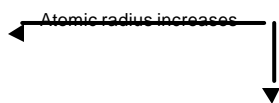
The Cl-Cl is 1.99 Å, implying a radius of 0.99 Å.

The C-Cl bond length in  $\text{CCl}_4$  is 1.77 Å, which is very close to the sum (0.77 + 0.99 Å) of the atomic radii for C and Cl.

## Periodic trends

**Atomic size** - The atomic radius increases going down a group. This can be explained by adding more and more energy level.

As one moves left to right within a period, the radii of the atoms decrease. As more and more protons are added to the nucleus, the nuclear charge of the elements increases. This has the effect of pulling the electrons closer to the nucleus and reducing the size of the atom.



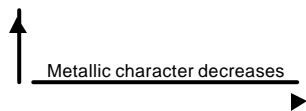
1 IA	2 IIA	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA
Li 0.152	Be 0.111	B 0.080	C 0.077	N 0.070	O 0.066	F 0.064
Na 0.186	Mg 0.160	Al 0.143	Si 0.118	P 0.111	S 0.104	Cl 0.099
K 0.227	Ca 0.197	Ga 0.122	Ge 0.125	As 0.116	Se 0.116	Br 0.115
Rb 0.248	Sr 0.215	In 0.163	Sn 0.141	Sb 0.145	Te 0.145	I 0.135
Cs 0.266	Ba 0.217	Tl 0.170	Pb 0.175	Bi 0.135		

Transition elements show many exceptions to the general trends.

## Metallic character

A metal reacts by losing one or more of its outermost electrons. As one moves up a group, the outermost electrons are closer to the nucleus. If the electrons and nucleus are closer together, it is more difficult for an atom to lose an electron.

Thus, the metallic character decreases from left to right across a period and bottom to top within a group.



## Ionization Energy

The ease with which electrons can be removed from an atom is an important indicator of the atom's chemical behavior.

The amount of energy necessary to remove an electron from a neutral atom in the gaseous state is called ionization energy. The first ionization energy  $I_1$ , is the energy needed to remove the first electron from a neutral atom.

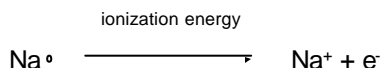


TABLE 7.2 Successive Values of Ionization Energies,  $I_n$ , for the Elements Sodium Through Argon (kJ/mol)

Element	$I_1$	$I_2$	$I_3$	$I_4$	$I_5$	$I_6$	$I_7$
Na	496	4560					
Mg	738	1450	7730				
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

The trend ( $I_1 < I_2 < I_3$ ,) arises because the positive nuclear charge that provides the attractive force remains the same, whereas the number of electrons, which produce repulsive interactions, steadily increases. As a result, the effective nuclear charge experienced by the remaining electrons increases.

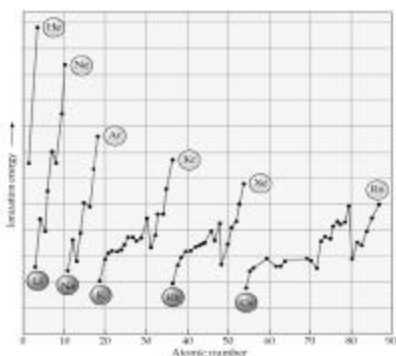
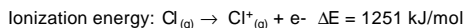
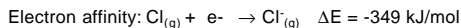


TABLE 7.3 Characteristic Properties of Metals and Nonmetals

Metals	Nonmetals
Have a shiny luster; various colors, although most are silvery.	Do not have a luster; various colors
Solids are malleable and ductile	Solids are usually brittle; some are hard, and some are soft
Good conductors of heat and electricity	Poor conductors of heat and electricity
Most metal oxides are ionic solids that are basic	Most nonmetallic oxides are molecular substances that form acidic solutions
Tend to form cations in aqueous solution	Tend to form anions or oxyanions in aqueous solution

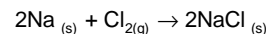
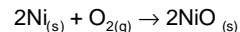
## ELECTRON AFFINITIES

The energy change that occurs when an electron is added to a gaseous atom is called the electron affinity because it measures the attraction, or affinity, of the atom for the added electron.



## METALS, NONMETALS, AND METALLOIDS

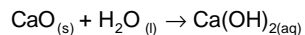
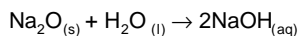
Compounds of metals with nonmetals tend to be ionic substances. Most metal oxides and halides are ionic solids.



The oxides are particularly important because of the great abundance of oxygen in our environment.

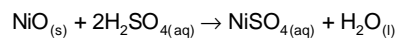
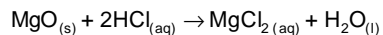
Most metal oxides are basic oxides; those that dissolve in water react to form metal hydroxides.

Metal oxide + water  $\rightarrow$  metal hydroxide



Metal oxides also demonstrate their basicity by reacting with acids to form salts and water.

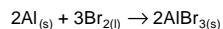
Metal oxide + acid  $\rightarrow$  salt + water



## NONMETALS

Nonmetals, in reacting with metals, tend to gain electrons and become anions.

Metal + nonmetal  $\rightarrow$  salt



The nonmetals commonly gain enough electrons to fill their outer  $p$  subshell completely, giving a noble-gas electron configuration.

Compounds composed entirely of nonmetals are molecular substances. For example, the oxides ( $\text{CO}_2$ ), halides ( $\text{CCl}_4$ ), and hydrides ( $\text{NH}_3$ ) of the nonmetals are molecular substances that tend to be gases, liquids, or low-melting solids.

Most nonmetal oxides are acidic oxides; those that dissolve in water react to form acids.

Nonmetal oxide + water  $\rightarrow$  acid

