8.4: Atomic Properties and Chemical Reactivity

Learning Objectives

- To understand the basic properties separating Metals from Nonmetals and Metalloids

An element is the simplest form of matter that cannot be split into simpler substances or built from simpler substances by any ordinary chemical or physical method. There are 118 elements known to us, out of which 92 are naturally occurring, while the rest have been prepared artificially. Elements are further classified into metals, non-metals, and metalloids based on their properties, which are correlated with their placement in the periodic table.

Table 8.4.1: Characteristic properties of metallic and non-metallic elements:

Metallic Elements	Nonmetallic elements
Distinguishing luster (shine)	Non-lustrous, various colors
Malleable and ductile (flexible) as solids	Brittle, hard or soft
Conduct heat and electricity	Poor conductors
Metallic oxides are basic, ionic	Nonmetallic oxides are acidic, covalent
Form cations in aqueous solution	Form anions, oxyanions in aqueous solution

Metals

With the exception of hydrogen, all elements that form positive ions by losing electrons during chemical reactions are called metals. Thus metals are electropositive elements with relatively low ionization energies. They are characterized by bright luster, hardness, ability to resonate sound and are excellent conductors of heat and electricity. Metals are solids under normal conditions except for Mercury.

Physical Properties of Metals

Metals are lustrous, malleable, ductile, good conductors of heat and electricity. Other properties include:

- **State**: Metals are solids at room temperature with the exception of mercury, which is liquid at room temperature (Gallium is liquid on hot days).
- **Luster**: Metals have the quality of reflecting light from their surface and can be polished e.g., gold, silver and copper.
- **Malleability:** Metals have the ability to withstand hammering and can be made into thin sheets known as foils. For example, a sugar cube sized chunk of gold can be pounded into a thin sheet that will cover a football field.
- **Ductility:** Metals can be drawn into wires. For example, 100 g of silver can be drawn into a thin wire about 200 meters long.
- **Hardness:** All metals are hard except sodium and potassium, which are soft and can be cut with a knife.
- **Valency:** Metals typically have 1 to 3 electrons in the outermost shell of their atoms.
- **Conduction**: Metals are good conductors because they have free electrons. Silver and copper are the two best conductors of heat and electricity. Lead is the poorest conductor of heat. Bismuth, mercury and iron are also poor conductors
- **Density**: Metals have high density and are very heavy. Iridium and osmium have the highest densities whereas lithium has the lowest density.
- **Melting and Boiling Points**: Metals have high melting and boiling points. Tungsten has the highest melting and boiling points whereas mercury has the lowest. Sodium and potassium also have low melting points.

Chemical Properties of Metals

Metals are electropositive elements that generally form *basic* or *amphoteric* oxides with oxygen. Other chemical properties include:

- **Electropositive Character**: Metals tend to have low ionization energies, and *typically lose electrons (i.e. are* ***oxidized****) when they undergo chemical reactions* They normally do not accept electrons. For example:
    - Alkali metals are always 1^+ (lose the electron in *s* subshell)
    - Alkaline earth metals are always 2^+ (lose both electrons in *s* subshell)

- Transition metal ions do not follow an obvious pattern, 2^+ is common (lose both electrons in *s* subshell), and 1^+ and 3^+ are also observed

$$\mathrm{Na}^0 \rightarrow \mathrm{Na}^+ + \mathrm{e}^- \tag{8.4.1}$$

$$\mathrm{Mg}^0 \rightarrow \mathrm{Mg}^{2+} + 2\,\mathrm{e}^- \tag{8.4.2}$$

$$\mathrm{Al}^0 \rightarrow \mathrm{Al}^{3+} + 3\,\mathrm{e}^- \tag{8.4.3}$$

Compounds of metals with non-metals tend to be ***ionic*** in nature. Most metal oxides are basic oxides and dissolve in water to form ***metal hydroxides***:

$$\mathrm{Na_2O(s)} + \mathrm{H_2O(l)} \rightarrow 2\,\mathrm{NaOH(aq)} \tag{8.4.4}$$

$$\mathrm{CaO(s)} + \mathrm{H_2O(l)} \rightarrow \mathrm{Ca(OH)_2(aq)} \tag{8.4.5}$$

Metal oxides exhibit their ***basic*** chemical nature by reacting with ***acids*** to form metal ***salts*** and water:

$$\mathrm{MgO(s)} + \mathrm{HCl(aq)} \rightarrow \mathrm{MgCl_2(aq)} + \mathrm{H_2O(l)} \tag{8.4.6}$$

$$\mathrm{NiO(s)} + \mathrm{H_2SO_4(aq)} \rightarrow \mathrm{NiSO_4(aq)} + \mathrm{H_2O(l)} \tag{8.4.7}$$

Example 8.4.1

What is the chemical formula for aluminum oxide?

Solution

Al has a 3+ charge, the oxide ion is O^{2-}, thus Al_2O_3.

Example 8.4.2

Would you expect it to be solid, liquid or gas at room temperature?

Solutions

Oxides of metals are characteristically solid at room temperature

Example 8.4.3

Write the balanced chemical equation for the reaction of aluminum oxide with nitric acid:

Solution

Metal oxide + acid -> salt + water

$$\mathrm{Al_2O_3(s)} + 6\,\mathrm{HNO_3(aq)} \rightarrow 2\,\mathrm{Al(NO_3)_3(aq)} + 3\,\mathrm{H_2O(l)}$$

Nonmetals

Elements that tend to gain electrons to form anions during chemical reactions are called non-metals. These are electronegative elements with high ionization energies. They are non-lustrous, brittle and poor conductors of heat and electricity (except graphite). Non-metals can be gases, liquids or solids.

Physical Properties of Nonmetals

- **Physical State**: Most of the non-metals exist in two of the three states of matter at room temperature: gases (oxygen) and solids (carbon). Only bromine exists as a liquid at room temperature.
- **Non-Malleable and Ductile**: Non-metals are very brittle, and cannot be rolled into wires or pounded into sheets.
- **Conduction**: They are poor conductors of heat and electricity.
- **Luster:** These have no metallic luster and do not reflect light.
- **Melting and Boiling Points**: The melting points of non-metals are *generally* lower than metals, but are highly variable.
- Seven non-metals exist under standard conditions as *diatomic molecules*: $\mathrm{H_2(g)}$, $\mathrm{N_2(g)}$, $\mathrm{O_2(g)}$, $\mathrm{F_2(g)}$, $\mathrm{Cl_2(g)}$, $\mathrm{Br_2(l)}$, $\mathrm{I_2(s)}$.

Chemical Properties of Nonmetals

Non-metals have a tendency to gain or share electrons with other atoms. They are electronegative in character. Nonmetals, when reacting with metals, tend to gain electrons (typically *attaining noble gas electron configuration)* and become ***anions:***

$$3\,Br_2(l) + 2\,Al(s) \rightarrow 2\,AlBr_3(s)$$

Compounds composed entirely of nonmetals are covalent substances. They generally form acidic or neutral oxides with oxygen that that dissolve in water to form acids:

$$CO_2(g) + H_2O(l) \rightarrow \underbrace{H_2CO_3(aq)}_{\text{carbonic acid}}$$

As you may know, carbonated water is slightly acidic (carbonic acid).

Nonmetal oxides can combine with bases to form salts.

$$CO_2(g) + 2\,NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$$

Metalloids

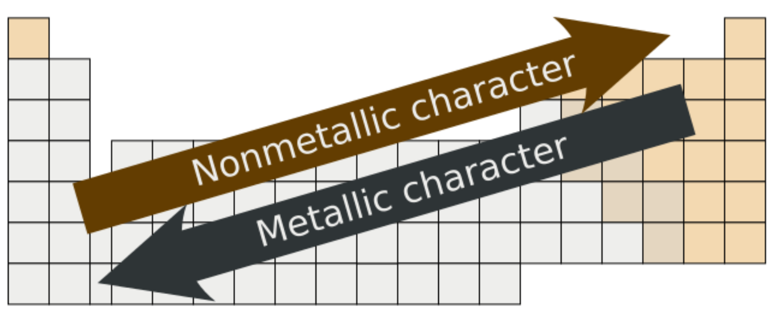
Metalloids have properties intermediate between the metals and nonmetals. Metalloids are useful in the semiconductor industry. Metalloids are all solid at room temperature. They can form alloys with other metals. Some metalloids, such as silicon and germanium, can act as electrical conductors under the right conditions, thus they are called semiconductors. Silicon for example appears lustrous, but is *not* malleable nor ductile (it is *brittle* - a characteristic of some nonmetals). It is a much poorer conductor of heat and electricity than the metals. The physical properties of metalloids tend to be metallic, but their chemical properties tend to be non-metallic. The oxidation number of an element in this group can range from +5 to -2, depending on the group in which it is located.

Table 8.4.2: Elements categorized into metals, non-metals and metalloids.

Metals	**Non-metals**	**Metalloids**
Gold	Oxygen	Silicon
Silver	Carbon	Boron
Copper	Hydrogen	Arsenic
Iron	Nitrogen	Antimony
Mercury	Sulfur	Germanium
Zinc	Phosphorus	

Trends in Metallic and Nonmetallic Character

Metallic character is strongest for the elements in the leftmost part of the periodic table, and tends to decrease as we move to the right in any period (nonmetallic character increases with increasing electronegativity and ionization energy values). Within any group of elements (columns), the metallic character increases from top to bottom (the electronegativity and ionization energy values generally decrease as we move down a group). This general trend is not necessarily observed with the transition metals.



Non metallic character increases bottom to top and left to right with the most nonmetal element in the top right corner. metallic character shows the opposite trend with the most metallic element being in the bottom left.

Contributors and Attributions

- Mike Blaber (Florida State University)
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The elements within the same group of the periodic table tend to exhibit similar physical and chemical properties. Four major factors affect reactivity of metals: nuclear charge, atomic radius, shielding effect and sublevel arrangement (of electrons). Metal reactivity relates to ability to lose electrons (oxidize), form basic hydroxides, form ionic compounds with non-metals. In general, the bigger the atom, the greater the ability to lose electrons. The greater the shielding, the greater the ability to lose electrons. Therefore, metallic character increases going down the table, and decreases going across -- so the most active metal is towards the left and down.

Group 1: The Alkali Metals

The word "alkali" is derived from an Arabic word meaning "ashes". Many sodium and potassium compounds were isolated from wood ashes (Na_2CO_3 and K_2CO_3 are still occasionally referred to as "soda ash" and "potash"). In the alkali group, as we go down the group we have elements Lithium (Li), Sodium (Na), Potassium (K), Rubidium (Rb), Cesium (Cs) and Francium (Fr). Several physical properties of these elements are compared in Table 8.4.1. These elements have all only one electron in their outermost shells. All the elements show metallic properties and have valence +1, hence they give up electron easily.

Table 8.4.1: General Properties of Group I Metals

Element	Electronic Configuration	Melting Point (°C)	Density (g/cm^3)	Atomic Radius	Ionization Energy (kJ/mol)
Lithium	$[He]2s^1$	181	0.53	1.52	520
Sodium	$[Ne]3s^1$	98	0.97	1.86	496
Potassium	$[Ar]4s^1$	63	0.86	2.27	419
Rubidium	$[Kr]5s^1$	39	1.53	2.47	403
Cesium	$[Xe]6s^1$	28	1.88	2.65	376

As we move down the group (from Li to Fr), the following trends are observed (Table 8.4.1):

- All have a single electron in an 's' valence orbital
- The melting point decreases
- The density increases
- The atomic radius increases
- The ionization energy decreases (first ionization energy)

> *The alkali metals have the lowest I_1 values of the elements*

This represents the relative ease with which the lone electron in the outer 's' orbital can be removed.

The alkali metals are very reactive, readily losing 1 electron to form an ion with a 1+ charge:

$$M \rightarrow M^+ + e-$$

Due to this reactivity, the alkali metals are found in nature only as compounds. The alkali metals combine directly with most nonmetals:

- React with hydrogen to form solid *hydrides*

$$2M_{(s)} + H_{2(g)} \rightarrow 2MH(s)$$

(Note: hydrogen is present in the metal hydride as the ***hydride*** H^- ion)

- React with sulfur to form solid *sulfides*

$$2M_{(s)} + S_{(s)} \rightarrow M_2S_{(s)}$$

React with chlorine to form solid *chlorides*

$$2M_{(s)} + Cl_{2(g)} \rightarrow 2MCl_{(s)}$$

Alkali metals react with water to produce hydrogen gas and alkali metal hydroxides; this is a very exothermic reaction (Figure 8.4.1).

$$2M_{(s)} + 2H_2O_{(l)} \rightarrow 2MOH_{(aq)} + H_{2(g)}$$

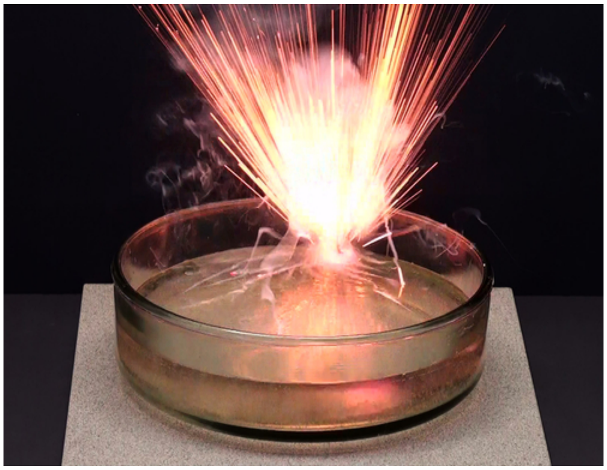

Figure 8.4.1: A small piece of potassium metal explodes as it reacts with water. (CC SA-BY 3.0; Tavoromann)

The reaction between alkali metals and oxygen is more complex:

- A common reaction is to form metal oxides which contain the ***O^{2-} ion***

$$4Li_{(s)} + O_{2(g)} \rightarrow \underbrace{2Li_2O_{(s)}}_{\text{lithium oxide}}$$

Other alkali metals ***can*** form metal peroxides (contains O_2^{2-} ion)

$$2Na(s) + O_{2(g)} \rightarrow \underbrace{Na_2O_{2(s)}}_{\text{sodium peroxide}}$$

K, Rb and Cs ***can*** also form superoxides (O_2^- ion)

$$K(s) + O_{2(g)} \rightarrow \underbrace{KO_{2(s)}}_{\text{potassium superoxide}}$$

Colors via Absorption

The color of a chemical is produced when a valence electron in an atom is excited from one energy level to another by visible radiation. In this case, the particular frequency of light that excites the electron is ***absorbed***. Thus, the remaining light that you see is white light devoid of one or more wavelengths (thus appearing colored). Alkali metals, having lost their outermost electrons, have no electrons that can be excited by visible radiation. Alkali metal salts and their aqueous solution are colorless unless they contain a colored anion.

Colors via Emission

When alkali metals are placed in a flame the ions are reduced (gain an electron) in the lower part of the flame. The electron is excited (jumps to a higher orbital) by the high temperature of the flame. When the excited electron falls back down to a lower orbital a photon is released. The transition of the valence electron of sodium from the 3p down to the 3s subshell results in release of a photon with a wavelength of 589 nm (yellow)

Flame colors:

- Lithium: crimson red
- Sodium: yellow
- Potassium: lilac

Group 2: The Alkaline Earth Metals

Compared with the alkali metals, the alkaline earth metals are typically harder, more dense, melt at a higher temperature. The first ionization energies (I_1) of the alkaline earth metals are not as low as the alkali metals. The alkaline earth metals are therefore less reactive than the alkali metals (Be and Mg are the least reactive of the alkaline earth metals). Several physical properties of these elements are compared in Table $8.4.2$.

Table $8.4.2$: General Properties of Group 2 Metals

Element	Electronic Configuration	Melting Point (°C)	Density (g/cm^3)	Atomic Radius	Ionization Energy (kJ/mol)
Beryllium	$[He]2s^2$	1278	1.85	1.52	899
Magnesium	$[Ne]3s^2$	649	1.74	1.60	738
Calcium	$[Ar]4s^2$	839	1.54	1.97	590
Strontium	$[Kr]5s^2$	769	2.54	2.15	549
Barium	$[Xe]6s^2$	725	3.51	2.17	503

Calcium, and elements below it, react readily with water at room temperature:

$$Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$$

The tendency of the alkaline earths to lose their two valence electrons is demonstrated in the reactivity of Mg towards chlorine gas and oxygen:

$$Mg_{(s)} + Cl_{2(g)} \rightarrow MgCl_{2(s)}$$

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

The 2+ ions of the alkaline earth metals have a noble gas like electron configuration and are thus form colorless or white compounds (unless the anion is itself colored). Flame colors:

- Calcium: brick red
- Strontium: crimson red
- Barium: green

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Learning Objectives

- To gain a descriptive understanding of the chemical properties of Hydrogen, the group 16, 17 and 18 elements.

Non-metallic character is the ability to be reduced (be an oxidizing agent), form acidic hydroxides, form covalent compounds with non-metals. These characteristics increase with a larger nuclear charge and smaller radius, with no increase in shielding. The most active non-metal would be the one farthest up and to the right -- not including the noble gases (non-reactive.)

Hydrogen

Hydrogen has a 1s^1 electron configuration and is placed above the alkali metal group. Hydrogen is a ***non-metal***, which occurs as a ***gas*** (H_2) under normal conditions.

- Its ionization energy is considerably higher (due to lack of shielding, and thus higher Z_{eff}) than the rest of the Group 1 metals and is more like the ***nonmetals***

- Hydrogen generally reacts with other nonmetals to form molecular compounds (typically highly exothermic)
- Hydrogen reacts with active metals to form metal hydrides which contain the ***H^- hydride ion***:

$$2Na_{(s)} + H_{2(g)} \rightarrow 2NaH_{(s)} \tag{8.4.8}$$

- Hydrogen can also lose an electron to yield the aqueous $H^+_{(aq)}$ hydronium ion.

Group 16: The Oxygen Family

As we proceed down group 16 the elements become more metallic in nature:

- Oxygen is a gas, the rest are solids
- Oxygen, sulfur and selenium are nonmetals
- Tellurium is a metalloid with some metal properties
- Polonium is a metal

Oxygen can be found in two molecular forms, O_2 and O_3 (ozone). These two forms of oxygen are called ***allotropes*** (*different forms of the same element in the same state*)

$$3O_{2(g)} \rightarrow 2O_{3(g)} \quad \Delta H = 284.6 \ kJ/mol \tag{8.4.9}$$

the reaction is endothermic, thus ozone is less stable that O_2

> Oxygen has a great tendency to attract electrons from other elements (i.e. to "oxidize" them)

- Oxygen in combination with metals is almost always present as the O^{2-} ion (which has noble gas electronic configuration and is particularly stable)
- Two other oxygen anions are observed: ***peroxide (O_2^{2-}) and superoxide (O_2^-)***

Sulfur

Sulfur also exists in several allotropic forms, the most common stable allotrope is the yellow solid S_8 (an 8 member ring of sulfur atoms). Like oxygen, ***sulfur has a tendency to gain electrons*** from other elements, and to form ***sulfides*** (which contain the S^{2-} ion). This is particular true for the active metals:

$$16Na_{(s)} + S_{8(s)} \rightarrow 8Na_2S_{(s)} \tag{8.4.10}$$

Note: most sulfur in nature is present as a metal-sulfur compound. Sulfur chemistry is more complex than that of oxygen.

Group 17: The Halogens

"Halogen" is derived from the Greek meaning "salt formers"

- Astatine is radioactive and rare, and some of its properties are unknown
- All the halogens are ***nonmetals***
- Each element consists of diatomic molecules under standard conditions

Colors of diatomic halogens: (*not* flame colors)

- Fluorine: pale yellow
- Chlorine: yellow green
- Bromine: reddish brown
- Iodine: violet vapor

The halogens have some of the most negative electron affinities (i.e. large exothermic process in gaining an electron from another element)

$$X_2 + 2e^- \rightarrow 2X^- \tag{8.4.11}$$

- Fluorine and chlorine are the most reactive halogens (highest electron affinities). Fluorine will remove electrons from almost any substance (including several of the noble gases from Group 18).

Note

The chemistry of the halogens is dominated by their tendency to gain electrons from other elements (forming a halide ion)

In 1992, **22.3 billion pounds** of chlorine was produced. Both chlorine and sodium can be produced by electrolysis of molten sodium chloride (table salt). The electricity is used to strip electrons from chloride ions and transfer them to sodium ions to produce chlorine gas and solid sodium metal

Chlorine reacts slowly with water to produce hydrochloric acid and hypochlorous acid:

$$Cl_{2(g)} + H_2O_{(l)} \rightarrow HCl_{(aq)} + HOCl_{(aq)}.5 \tag{8.4.12}$$

Hypochlorous acid is a disinfectant, thus chlorine is a useful addition to swimming pool water

The halogens react with most metals to form ionic halides:

$$Cl_{2(g)} + 2Na_{(s)} \rightarrow 2NaCl_{(s)} \tag{8.4.13}$$

Group 18: The Noble Gases

- Nonmetals
- Gases at room temperature
- monoatomic
- completely filled 's' and 'p' subshells
- large first ionization energy, ***but this decreases somewhat as we move down the group***

Rn is highly radioactive and some of its properties are unknown

They are exceptionally unreactive. It was reasoned that if any of these were reactive, they would most likely be Rn, Xe or Kr where the first ionization energies were lower.

Note

In order to react, they would have to be combined with an element which had a high tendency to remove electrons from other atoms. Such as fluorine.

Compounds of noble gases to date:

$$XeF_2 \ XeF_4 \ XeF_6$$

only one compound with Kr has been made

$$KrF_2$$

No compounds observed with He, Ne, or Ar; they are truly inert gases.

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Hydrogen

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- Its ionization energy is considerably higher (due to lack of shielding, and thus higher Z_{eff}) than the rest of the Group 1 metals and is more like the ***nonmetals***
- Hydrogen generally reacts with other nonmetals to form molecular compounds (typically highly exothermic)
- Hydrogen reacts with active metals to form metal hydrides which contain the ***H^- hydride ion***:

$$2Na_{(s)} + H_{2(g)} \rightarrow 2NaH_{(s)} \label{7.8.1}$$

- Hydrogen can also lose an electron to yield the aqueous $H^+_{(aq)}$ hydronium ion.

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