Periodic Trends

Elements within a period or group of the periodic table often show trends in physical and chemical properties. The variation in relative sizes of the halogens is shown above.

DEPARTMENT OF CHEMISTRY

UNIVERSITY OF KANSAS
Periodic Trends

Introduction

In the modern periodic table (shown below in Figure 1), elements are arranged according to increasing atomic number in horizontal rows called “periods.” In Figure 1, atomic numbers, which represent the number of protons in an atom of a given element, are listed directly above the element symbols.

The structure of the periodic table is such that elements with similar properties are aligned vertically in columns called “groups” or “families.” As indicated in Figure 1, each group has a number (1-18) associated with it. Select groups have also been assigned special names. For instance, the elements in Group 1 (hydrogen excluded) are called the alkali metals. These metallic elements react with oxygen to form bases. They also form alkaline (basic) solutions when mixed with water.

Elements in other columns of the periodic table have also been given special “family” names. For instance, the elements in Groups 2, 11 (Cu, Ag, and Au), 16, 17, and 18 are commonly referred to as the alkaline earth metals, the coinage metals, the chalcogens, the halogens, and the noble gases, respectively. Many of these names reflect, unsurprisingly, the reactivity (or the lack thereof) of the elements found within the given group.

The arrangement of the elements in the periodic table gives rise to smoothly varying trends in properties as one moves down the periodic table within a specific group or as one moves horizontally along a given row (or period) in the periodic table. For instance, atomic radius values decrease across a given period of the periodic table and increase down a given group. For atomic radius, the trend always holds. For other properties, there may be exceptions in the midst of the overall behavior. One example is ionization energy, which is the energy required to remove one
electron from a neutral atom of an element in its gaseous state. Overall, the ionization energy increases from left to right across a period. If we look closely, though, we can see that in row two, this does not hold for Be to B, and again between N and O. Recall from Chapter 4 of your Burdge and Overby text the reasons why.

In Part 1 of this experiment, you will investigate the relative solubilities of select alkaline earth compounds and use your observations to rank the Group 2 metals (Ba, Ca, Mg, and Sr) according to their tendency to form insoluble ionic compounds. The final exercise in this experiment will involve studying the reactivity of select metals (Ca, Cu, Fe, Mg, Sn, and Zn) in order to generate an activity series for these elements. In each Part, the observed trends will be compared with the arrangement of the elements in the periodic table.

Part 1 Background – Relative Solubilities of Select Alkaline Earth Compounds

The alkaline earth elements are reactive metals that tend to lose electrons (i.e., undergo oxidation) and form +2 cations in their ionic compounds. By losing two electrons, these elements attain the same number of electrons as noble gases, which have particularly stable numbers and arrangements of electrons. For instance, calcium loses two electrons to obtain the same number of electrons as argon. The resulting cations can form ionic compounds with various anions, and the resulting ionic compounds are soluble in aqueous environments to different extents.

The first part of this experiment involves examining the relative solubilities of select Group 2 ionic compounds and noting any observed periodic trends or recurring patterns in solubility based on the placement of the alkaline earth elements in the periodic table. Specifically, soluble nitrate solutions containing the Ba$^{2+}$, Ca$^{2+}$, Mg$^{2+}$, and Sr$^{2+}$ ions will be mixed one by one with aqueous solutions of H$_2$SO$_4$, Na$_2$CO$_3$, (NH$_4$)$_2$C$_2$O$_4$, and KIO$_3$. In each case, the appearance of a precipitate (i.e., an insoluble solid) will be noted. A detailed description of the precipitate will also be recorded. The observed solubility should vary somewhat smoothly among the alkaline earth elements. By observing these trends in solubility you should be able to confirm the ordering of the elements within this group.

Part 2 Background – Metal Activity Series and Chemical Equations

Most elements are classified as metals (see Figure 1). Some metals are highly reactive and readily oxidize (e.g., the alkali and alkaline earth metals), while others are resistant to oxidation and corrosion (e.g., gold and platinum). Reactive metals can participate in a number of reactions. Recall from section 4.4 of your book that we represent chemical reactions by placing the starting materials (reactants) on the left side of an arrow and the ending materials (products) on the right. For example, some metals react with water to form hydrogen gas, and we might write:

$$\text{Mg} \text{(s)} + 2 \text{HCl \text{(aq)}} \rightarrow \text{MgCl}_2 \text{(aq)} + \text{H}_2 \text{(g)}$$

which shows that we start with magnesium and hydrochloric acid. As the reaction proceeds, atoms and their electrons rearrange to produce magnesium chloride salt and molecular hydrogen. The subscripts indicate that the starting magnesium is a solid metal, the hydrochloric acid and magnesium chloride are both dissolved (soluble) in water, and that the molecular hydrogen is a gas. As a gaseous product, the hydrogen should form bubbles in the solution. If we were to see an insoluble precipitate form during the reaction, what subscript would that get?

Reactions between metals and acids can also liberate this highly flammable, odorless diatomic gas. Metal displacement reactions are another type of oxidation-reduction process involving metals. In metal displacement reactions, a more active metal displaces (or replaces) a less active metal from its compound. Specific examples of some common reactions involving metals are given below.
3a—Metals reacting with oxygen

Magnesium metal reacts with oxygen to form solid magnesium oxide.

\[ 2 \text{Mg (s)} + \text{O}_2 (g) \rightarrow 2 \text{MgO (s)} \]  

(3)

As shown in Figure 2, this combustion reaction gives off a significant amount of visible light. Ultraviolet radiation is also emitted. This is an example of an oxidation-reduction (or redox) reaction. In this case, magnesium metal is oxidized by oxygen.

3b—Metals reacting with water

Group 2 metals also react with water to form hydrogen gas. These reactions are similar to (but generally less vigorous than) those involving alkali metals and water. Figure 3 shows what happens when a piece of calcium metal is immersed in water. The “bubbles” in this picture indicate the evolution of a gas, in this case hydrogen.

\[ \text{Ca (s)} + 2 \text{H}_2\text{O (l)} \rightarrow \text{Ca(OH)}_2 (aq) + \text{H}_2 (g) \]  

(4)

This reaction is an example of a hydrogen displacement reaction since calcium metal displaces hydrogen from water.

3c—Metals reacting with acid

Hydrogen gas is generated when aluminum metal is added to a solution of hydrochloric acid.

\[ 2 \text{Al (s)} + 6 \text{HCl (aq)} \rightarrow 2 \text{AlCl}_3 (aq) + 3 \text{H}_2 (g) \]  

(5)

This is another example of a hydrogen displacement redox reaction. In this case, aluminum undergoes oxidation and displaces hydrogen from hydrochloric acid. Many metals are capable of displacing hydrogen from acids. Some metals that react slowly (or not at all) with water react readily with acids to produce hydrogen gas. This is true for aluminum. This is also true for magnesium metal. Mg, which does not react in cold water, reacts quite readily with hydrochloric acid to form H₂ gas. This reaction is shown below.

\[ \text{Mg (s)} + 2 \text{HCl (aq)} \rightarrow \text{MgCl}_2 (aq) + \text{H}_2 (g) \]  

(6)

Keep in mind that not all metals react with water or acids to release hydrogen gas. If copper readily reacted with water to form hydrogen gas, household plumbing would certainly not be constructed out of this metal. This example illustrates how the reactivity of a metal strongly influences its application. Jewelry, for instance, would not be fashioned out of gold or platinum if these metals were highly reactive and prone to corrosion.
3d—Metal displacement

When aluminum metal is added to a solution of copper(II) chloride, the aluminum displaces the copper from the chloride compound to form copper metal.

\[ 2\ Al_{(s)} + 3\ CuCl_{2(aq)} \rightarrow 2\ AlCl_{3(aq)} + 3\ Cu_{(s)} \]  \hspace{1cm} (7)

This is an example of a metal displacement redox reaction. In these types of reactions a free metal (i.e., one in the elemental state) displaces (or replaces) another metal from its compound. The balanced net ionic equation for this process is shown below.

\[ 2\ Al_{(s)} + 3\ Cu^{2+}_{(aq)} \rightarrow 2\ Al^{3+}_{(aq)} + 3\ Cu_{(s)} \]  \hspace{1cm} (8)

Examination of the net ionic equation reveals that aluminum undergoes oxidation in this process (from Al to Al\(^{3+}\)), and copper is reduced from Cu\(^{2+}\) to Cu. This reaction occurs because aluminum metal is a stronger reducing agent (i.e., it is more easily oxidized) than copper metal. As a result, aluminum is said to be a more “active” metal than copper. This reaction demonstrates that a more active metal can displace a less active metal from its ionic compound. The opposite, however, is not true (at least not without adding sufficient energy to drive the nonspontaneous process). For instance, if a piece of copper were immersed in an aqueous AlCl\(_3\) solution, no spontaneous reaction would occur since copper is a less active metal than aluminum. If a piece of magnesium were immersed in this same AlCl\(_3\) solution, the following reaction would occur:

\[ 3\ Mg_{(s)} + 2\ AlCl_{3(aq)} \rightarrow 3\ MgCl_{2(aq)} + 2\ Al_{(s)} \]  \hspace{1cm} (9)

indicating that magnesium is a more active metal than aluminum.

The information given in the previous sections on metal reactivity can be used to develop a metal “activity series.” An activity series ranks elements in order of their reactivity. In the section titled “Metals reacting with water,” the Group 2 metals were ordered according to increasing reactivity (or increasing strength as a reducing agent / electron donor). This order was found to be: Mg < Ca < Sr < Ba. Hence, barium is the most active metal in this series. In the “Metals reacting with acids” section, it was revealed that magnesium and aluminum react with acids to form H\(_2\) gas but copper does not. This indicates that Mg and Al are more reactive and more easily oxidized than copper. In the “Metal displacement” section it was revealed that magnesium is a more active metal (i.e., a stronger reducing agent / better electron donor) than aluminum. It was also shown (once again) that aluminum is a more reactive metal than copper. These observations can be used to rank these three metals in order of increasing reactivity. The resulting order would be: Cu < Al < Mg. Combining this activity series with the aforementioned Group 2 activity series gives: Cu < Al < Mg < Ca < Sr < Ba. Part 2 of this experiment involves conducting a series of reactions with six different metals (Ca, Cu, Fe, Mg, Sn, and Zn) and using the acquired observational data to develop an activity series describing the relative reactivities of these metals.
Pre-lab Safety: Goggles must be worn at all times. This lab involves working with a number of chemicals including reactive metals. Care should be taken when transporting materials across the lab. Solutions and metals should be mixed together only as described in the laboratory procedure. Barium solutions should be handled with care as barium compounds are toxic. If a solution containing Ba\(^{2+}\) is spilled on your skin, rinse the affected area immediately. Hydrogen gas will be generated in some of the reactions carried out in this experiment. This colorless, odorless gas is highly flammable. Thus, at no time should there be an open flame in the lab. All solutions and solid wastes should be discarded in the appropriate waste container(s) as instructed by your TA. None of the materials used in this experiment should be disposed of down the drain.

Pre-lab Assignment: Please write out the following in your lab notebook and print required graphs; items may be stapled or otherwise bound. This assignment must be completed before the beginning of lab. You will not be allowed to start the experiment until this assignment has been completed and accepted by your TA.

1) You will construct and print two plots that demonstrate different periodic trends.
   a. Atomic radius:
      i. Go to the website www.ptable.com. Click on the "Properties" tab at the top of the screen. This will open up a periodic table with a list of various chemical and physical properties. Click the button for "Radius."
      ii. The table should show several options in the upper right corner, and by default, "Calculated Radius" is selected. Keep this option. For each element, the atomic number appears above the elemental symbol, while the radius (in pm) appears below.
      iii. Open Excel and start a new workbook.
      iv. In your Excel spreadsheet, record the atomic numbers of the Period 4 elements in one column. Type the atomic radii of the corresponding elements in the neighboring column. Use these data to create an "Atomic radius vs. Atomic Number" XY scatter plot. Choose the chart sub-type option that graphs data points connected by lines. The radii (in pm) should be plotted on the y-axis; atomic number should be plotted on the x-axis. Please talk to your TA if you have questions or concerns about using Excel or creating a graph in Excel. Your graph should have a title and both axes should be properly labeled. These labels should include units if appropriate.
      v. Add a second line to the title section with your name and the date.
   b. Ionization energy:
      i. Go to www.ptable.com and select the tab for "Properties.” Find and click the button for “Ionization.”
      ii. Similarly to a-ii) – a-v) above, construct a graph that shows the first ionization energy as a function of atomic number for the period 4 elements. Be sure to include units on your axes.
   c. Trends: Based on your two graphs, which property—atomic radius or ionization energy—follows a strict trend? Which property follows a general trend? Explain your reasoning.
2) For each alkaline earth metal used Part 2 of this experiment, determine the total number of electrons its cation would have; also provide the identity of the nearest noble gas on the periodic table.

3) Form a table of alkaline earth metal cations and the various anions listed in Part 2 of this experiment. For each compound, find information online: Is the resulting compound (cation + anion) soluble or insoluble in water?

4) Define the terms ionization energy and electron affinity. Define the terms oxidation and reduction. How are these sets of terms similar? How are they different?

5) The general reactivity series for the alkali metals (rather than the alkaline earth metals) is Li < Na < K < Cs < Rb.
   a. For each, find the value of the first ionization energy. How does the trend in ionization energy compare with the reactivity series for alkali metals?
   b. What do you anticipate will be the trend in reactivity for the alkaline earth metals? In words or pictures, provide your reasoning.

6) Write the chemical equation for this process: "Copper metal is placed in a solution of silver nitrate. This produces silver metal and copper nitrate solution." You do NOT need to balance the equation, which is discussed later in the semester.

In addition to these pre-lab requirements, a short quiz may be given at the beginning of lab based on the material in this lab write-up.

Procedural

Part 1 – Relative Solubilities of Select Alkaline Earth Compounds

In this part of the experiment, nitrate solutions of barium, calcium, magnesium, and strontium will be mixed with solutions containing the sulfate (SO$_4^{2-}$), carbonate (CO$_3^{2-}$), oxalate (C$_2$O$_4^{2-}$), and iodate (IO$_3^{-}$) polyatomic anions. Any changes that might indicate a chemical reaction (e.g., a persistent color change, the formation of a precipitate, or the evolution of a gas) should be noted and described in detail in your lab notebook. Please follow the steps given below.

1. Sketch the following table in your notebook. Leave sufficient room to record detailed observations.

<table>
<thead>
<tr>
<th>Group 2 nitrates</th>
<th>test reagents</th>
<th>H$_2$SO$_4$</th>
<th>Na$_2$CO$_3$</th>
<th>(NH$_4$)$_2$C$_2$O$_4$</th>
<th>KIO$_3$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ba(NO$_3$)$_2$</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ca(NO$_3$)$_2$</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>Mg(NO$_3$)$_2$</td>
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<td></td>
<td></td>
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<tr>
<td>Sr(NO$_3$)$_2$</td>
<td></td>
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<td></td>
</tr>
</tbody>
</table>
2. Test 1 - Reacting the Group 2 Nitrates with \( \text{H}_2\text{SO}_4 \)

In the first qualitative test, you will study the solubility of Group 2 sulfate salts. Obtain 4 small test tubes. Place 10 drops of 1 M \( \text{H}_2\text{SO}_4 \) in each test tube. Add 5 drops of 0.1 M \( \text{Ba(NO}_3\text{)}_2 \) to the first test tube. Record your observations. Repeat this procedure using the remaining Group 2 nitrate solutions (i.e., add 5 drops of 0.1 M \( \text{Ca(NO}_3\text{)}_2 \), 0.1 M \( \text{Mg(NO}_3\text{)}_2 \), and 0.1 M \( \text{Sr(NO}_3\text{)}_2 \) to the second, third, and fourth test tubes, respectively.) Record your observations. When you are finished, pour the test tube contents into the indicated waste container. Rinse all of the test tubes.

3. Test 2 - Reacting the Group 2 Nitrates with \( \text{Na}_2\text{CO}_3 \)

The second test reagent is sodium carbonate. Add 5 drops of 1 M \( \text{Na}_2\text{CO}_3 \) to four different test tubes. Add 5 drops of 0.1 M \( \text{Ba(NO}_3\text{)}_2 \) to the first test tube. Record your observations. Repeat this process with the other Group 2 nitrate solutions. In other words, add 5 drops of 0.1 M \( \text{Ca(NO}_3\text{)}_2 \), 0.1 M \( \text{Mg(NO}_3\text{)}_2 \), and 0.1 M \( \text{Sr(NO}_3\text{)}_2 \) to the second, third, and fourth test tubes, respectively, and record your observations. When you are finished, pour the test tube contents into the indicated waste container. Rinse all of the test tubes.

4. Test 3 - Reacting the Group 2 Nitrates with \( (\text{NH}_4)_2\text{C}_2\text{O}_4 \)

The third test reagent is ammonium oxalate. Add 5 drops of 0.1 M \( (\text{NH}_4)_2\text{C}_2\text{O}_4 \) to four different test tubes. Add 5 drops of 0.1 M \( \text{Ba(NO}_3\text{)}_2 \) to the first test tube. Record your observations. Repeat this process with the other Group 2 nitrate solutions. That is, add 5 drops of 0.1 M \( \text{Ca(NO}_3\text{)}_2 \), 0.1 M \( \text{Mg(NO}_3\text{)}_2 \), and 0.1 M \( \text{Sr(NO}_3\text{)}_2 \) to the second, third, and fourth test tubes, respectively. Record your observations in each case. When you are finished, pour the test tube contents into the indicated waste container. Rinse all of the test tubes.

5. Test 4 - Reacting the Group 2 Nitrates with \( \text{KIO}_3 \)

The last test reagent is potassium iodate. Add 5 drops of 0.1 M \( \text{KIO}_3 \) to four different test tubes. Add 5 drops of 0.1 M \( \text{Ba(NO}_3\text{)}_2 \) to the first test tube. Record your observations. Repeat this process with the other Group 2 nitrate solutions by adding 5 drops of 0.1 M \( \text{Ca(NO}_3\text{)}_2 \), 0.1 M \( \text{Mg(NO}_3\text{)}_2 \), and 0.1 M \( \text{Sr(NO}_3\text{)}_2 \) to the second, third, and fourth test tubes, respectively, and recording your observations. When you are finished, pour the test tube contents into the indicated waste container. Rinse all of the test tubes.

Discussion and Analysis: Use your observations to determine the combination of alkaline earth cations and test solution anions that produce a precipitate. Predict the chemical formulas of the observed precipitates. (Information from your prelab may be useful here.) Use your observations to identify any periodically recurring trends in the relative solubilities of the Group 2 sulfates, carbonates, oxalates, and iodates. The observed trends in solubility should vary somewhat smoothly among the alkaline earth elements. That is, they should either increase or decrease gradually through this group. By observing these trends in solubility you should be able to confirm the ordering of these elements in the periodic table.
Part 2 – Metal Activity Series

The reactions carried out in this part of the experiment will be used to generate an activity series involving the following six metals: Ca, Cu, Fe, Mg, Sn, and Zn. In the first test, a small piece of each of these metals will be added to separate test tubes containing hydrochloric acid. Any evidence of a reaction (or the absence thereof) will be noted. In the second battery of qualitative tests, each of the given metals will be placed in a variety of different metal ion solutions to determine if a metal displacement reaction occurs. Using the data obtained from these tests, you should be able to order the six metals in order of increasing reactivity (i.e., you should be able to rank them from least active to most active). Please follow the procedure outlined below.

Sketch the following table in your notebook. If you prefer, you can construct two separate tables. For instance, one table can be used to record the results of the tests between the various metals and hydrochloric acid. A second table can be used to record the observations that you made when you mixed the six different metals with the various solutions containing metal ions. Regardless of how many tables you construct in your lab notebook, make sure that you leave sufficient room to record detailed observations. Note: The shaded areas in the table shown below indicate tests that need not be (and will not be) performed as no displacement reaction will occur.

<table>
<thead>
<tr>
<th>metal</th>
<th>HCl</th>
<th>Ca(NO₃)₂</th>
<th>Cu(NO₃)₂</th>
<th>FeSO₄</th>
<th>Fe(NO₃)₃</th>
<th>Mg(NO₃)₂</th>
<th>SnCl₄</th>
<th>Zn(NO₃)₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ca</td>
<td></td>
<td>✅</td>
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<td></td>
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<tr>
<td>Cu</td>
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<td></td>
<td>✅</td>
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<tr>
<td>Fe</td>
<td></td>
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<td>✅</td>
<td>✅</td>
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<tr>
<td>Mg</td>
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<td>✅</td>
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<tr>
<td>Sn</td>
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<td>Zn</td>
<td></td>
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<td></td>
<td></td>
<td>✅</td>
</tr>
</tbody>
</table>

1. Metal Displacement Reactions

In this last series of tests, you will examine how each of the six metals react when added to solutions containing cations of the other metals. Begin by adding a small piece of calcium metal to six different test tubes. Next, add 5 drops of 0.1 M Cu(NO₃)₂ to the first test tube, 5 drops of 0.1 M FeSO₄ to the second test tube, 5 drops of 0.1 M Fe(NO₃)₃ to the third test tube, 5 drops of 0.1 M Mg(NO₃)₂ to the fourth test tube, 5 drops 0.1 M SnCl₄ to the fifth test tube, and 5 drops of 0.1 M Zn(NO₃)₂ to the sixth test tube. Record any changes that occur and note any indication of reaction, including color change, gas evolution, etc. Be sure to make careful observations and record detailed notes. If you observe a color change, document whether it was the color of the solution and/or the metal surface that changed. When you are finished, the test tube contents should be discarded in the appropriate waste container(s). Rinse all of the test tubes before moving on.
Repeat this general procedure for each of the remaining 5 metals. Keep in mind that each series will involve slightly different metal solutions as there is no need to react a given metal with a solution containing the same metal cation. For instance, you do not need to add copper to a solution of copper(II) nitrate. Likewise, you do not need to add iron metal to the iron(II) sulfate and iron(III) nitrate solutions. If you have questions about this procedure, please ask your TA before you begin mixing various metals and solutions. When you have finished this series of tests, you should have data recorded in every cell of your table (or tables) except for those shaded grey in the sample table. In a number of cases, there will be no reaction. A lack of reaction indicates that the metal that was added to the solution is less active than the metal in solution.

Discussion and Analysis: Use the data collected above to help you rank Ca, Cu, Fe, Mg, Sn, and Zn in terms of increasing reactivity. The approach used here should be similar to the one discussed in the introduction. You should, in fact, be able to rank most (if not all) of the metals by simply comparing the reactions that occurred when each metal was added to HCl. Clearly, those metals that reacted with HCl to produce hydrogen gas would be considered more active that those that exhibited no reaction. In addition, the metal that reacted the most vigorously in acid would be the most reactive metal. By continuing this line of reasoning you should be able to generate at least a preliminary activity series for the metals under study. Data collected in the metal displacement tests can be used to corroborate your findings and/or aid you in determining the correct ranking (or at least the one supported by your experimental data). In this case, the metal that exhibited no reaction when added to all of the different metal solutions would be considered the least active metal. Conversely, the metal that displaced all of the metals in solution would be considered the most active metal. You should be able to articulate in detail how you arrived at your final prediction regarding the relative reactivities of these metals. In other words, you must understand how your group came up with your metal activity series.


Glossary

**Chemical Equation**

A simplified representation of what occurs during a chemical reaction. By convention, starting materials (reactants) are placed on the left of an arrow. The ending materials are placed to the right of the arrow. The arrow itself represents the chemical change or rearrangements of electrons, atoms, etc. A “balanced” chemical equation obeys Conservation of Matter and Conservation of Mass.

**Reactant**

The material or chemical species present before a chemical reaction occurs.

**Reduction**

The gain of electrons; a decrease or reduction in oxidation number.

**Solubility**

The extent to which a substance dissolves in a liquid, typically water. Solubility can be quantitative, providing the mass or number of moles of the substance that can be dissolved in some volume or mass of liquid (e.g., g / L). Solubility can also be qualitative: “Soluble” indicates the material in a solution is dissolved, with the solution appearing clear or translucent. “Insoluble” indicates that the material in a solution is not dissolved, with the solution appearing cloudy from the formation of small solid particles; large enough solid particles may also sink to the bottom.

**Trend**

For a sequence of items, the variation, change, or development of some property in a specific direction, for example increasing or decreasing; in a strict trend, all adjacent pairs within the sequence obey the same direction of change, and, therefore, the same direction of change is maintained over the whole sequence; a general trend is one in which a dominant direction of change occurs over the sequence, with a majority of adjacent pairs within the sequence obeying the overall direction of change, and a minority of adjacent pairs showing the opposite direction of change.

**Group; family**

A collection of elements within the same column of a standard (chemical) periodic table of the elements.

**Oxidation**

The loss of electrons; also an increase in oxidation number.

**Period**

A collection of elements within the same row of a standard (chemical) periodic table of the elements.

**Periodic Property**

A property of an element that is predictable based on an element’s position in the periodic table.

**Reactivity**

The tendency of or extent to which a substance participates in chemical reaction or undergoes chemical change; the degree to which a material is reactive.

**Product**

The material or chemical species resulting from a chemical reaction.