

Periodic Trends

Introduction

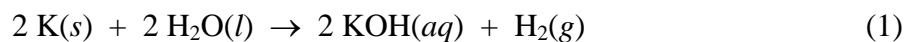
In the modern periodic table (shown below in Figure 1), elements are arranged according to increasing atomic number in horizontal rows called “periods.” In Figure 1, atomic numbers, which represent the number of protons in an atom of a given element, are listed directly above the element symbols.

1 H																	17 H	18 He
3 Li	4 Be											13 B	14 C	15 N	16 O	9 F	10 Ne	
11 Na	12 Mg	3	4	5	6	7	8	9	10	11	12	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar	
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr	
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe	
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn	
87 Fr	88 Ra	89 Ac	104 Rf	105 Ha	106 Sg	107 Ns	108 Hs	109 Mt										

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Figure 1. The modern periodic table. Elements in boxes shaded blue, orange, and purple are characterized as metals, metalloids, and nonmetals, respectively.

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. The equation shown below illustrates the reaction between potassium metal and water.



As shown in Figure 2c, the exothermicity of this reaction is sufficient to cause ignition of the hydrogen gas. The heat generated in this reaction is also high enough to melt the potassium metal. (The normal melting point of potassium is 63.4 °C.) While all alkali metals exhibit this



Figure 2. (a) Sodium stored in kerosene. (b) Potassium stored in an inert gas. (c) The reaction of potassium and water.

characteristic reactivity towards water, they do so to a varying degree. Indeed, the reactivity of these metals increases going down the group (*i.e.*, sodium is more reactive towards water than lithium; potassium is more reactive towards water than sodium, etc.). This ranking in reactivity ($\text{Li} < \text{Na} < \text{K}$) is in accordance with periodic law and can be predicted by periodic trends in first ionization energies, IE_1 , (see Figure 3) and the placement of these elements in the periodic table. Given the reactivity of alkali metals towards oxygen and water, it is no surprise that the elemental forms of these metals are not found free (*i.e.*, in an uncombined state) in nature. In fact, alkali metals are so reactive that once isolated they are typically stored in kerosene, mineral oil, or an inert gas (see Figures 2a and 2b) to prevent rapid oxidation.

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. The etymology of each of these names is quite interesting; however, many of them 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, IE_1 values *generally*¹

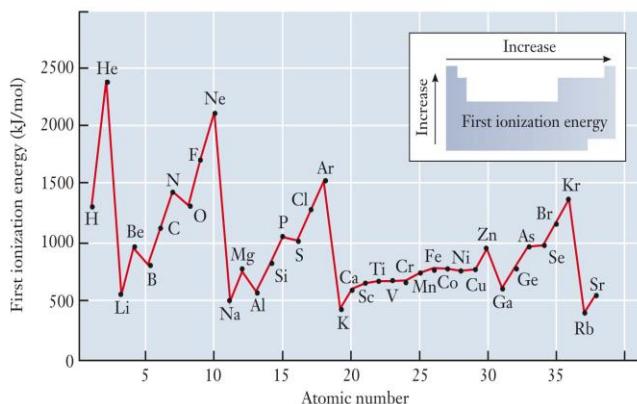


Figure 3. A plot of first ionization energies versus atomic number. Note that the alkali metals (Li, Na, K, and Rb) have the lowest IE_1 values in their respective periods. Furthermore, IE_1 values generally decrease going down a group in the periodic table.

¹The word “*generally*” is used here to highlight the fact that there are exceptions to this trend. Notable exceptions to this “*general*” trend are seen in Figure 3. See also page 170 in your text.

increase across a given period of the periodic table and *generally* decrease down a given group (see Figure 3). The reactivities of the elements and their compounds also follow well-defined trends within a group (as discussed previously in terms of the Group 1 metals) and across a given row. In Part 1 of this experiment, you will explore periodic trends in the melting point temperatures of metals. In Part 2, 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.

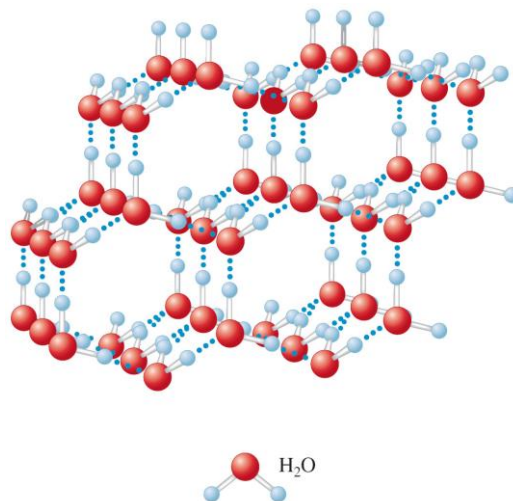


Figure 4. The hydrogen-bonded network in ice.

Part 1 – Periodic Trends in Melting Point Temperatures

Many properties of the elements show relatively well-defined periodic trends. Your textbook (*Chemistry for Engineering Students, 3/e* by Brown and Holme) highlights trends in several physical properties, including atomic radius, ionization energy, and electron affinity (see Figures 6.20, 6.21, 6.22, and 2.11 in your text and the associated discussions). There are, however, numerous other physical properties that also demonstrate periodic trends. In the first part of this laboratory exercise, you will explore the periodicity of melting point temperatures. This study will focus on the melting points of metals.

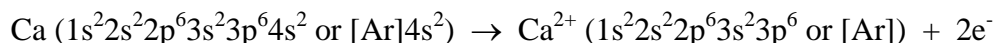
The melting point of a material is the temperature at which the solid and liquid forms of the substance are in equilibrium with one another at a given pressure. You are probably familiar with the fact that the melting point of ice (the solid form of water) is 0 °C (32 °F, 273.15 K) at 1 atm. The equilibrium process taking place at these conditions can be represented by the following chemical equation: $\text{H}_2\text{O}(s) \rightleftharpoons \text{H}_2\text{O}(l)$. This equation highlights the fact that this process involves a physical change versus a chemical change. While ice and liquid water certainly look very different, they are both composed of H_2O molecules.

The melting point of a substance can also be described as the temperature at which there is sufficient thermal energy to break or disrupt the cohesive bonds between neighboring particles. In the case of water, the solid to liquid phase transition involves disrupting the rigid, hydrogen-bonded network of water molecules within the ice structure (see Figure 4). Consequently, the temperature at which ice melts at a given pressure is intrinsically dependent on the strength of the intermolecular forces between the water molecules. Another way of stating this is that the melting point of a substance is dependent on (and a measure of) the cohesive forces between neighboring particles. Disrupting cohesive interactions requires energy. Hence, phase transitions require energy. Cohesive energies reflect how strongly each particle in a given system interacts with its neighbor. High energies are needed to overcome strong cohesive forces. Thus, the melting points of substances with strong cohesive forces between the neighboring atoms or

molecules are generally higher than those characterized by weak interactions between constituent particles. Since the melting point of a substance depends on the amount of energy needed to separate neighboring particles (such as the molecules of water in ice or the neighboring atoms in a metal), one would expect trends in the melting points of the elements to closely follow cohesive energy trends. Using data from the online site WebElements, you will investigate how the melting points of the elements vary with increasing atomic number along a given period or down a given group and use your understanding of electron configurations, bonding, cohesive energies, and packing (*i.e.*, crystal structure) to explain your observations.

Part 2 – 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 electron configuration as that of the preceding noble gas. For instance, the electron configuration of calcium is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ (or simply $[\text{Ar}]4s^2$). In ionic compounds, calcium atoms tend to lose their two 4s valence electrons to form the Ca^{2+} ion. This divalent cation has the same electron configuration as argon ($1s^2 2s^2 2p^6 3s^2 3p^6$ or $[\text{Ar}]$), the preceding noble gas. The oxidation half-reaction shown below explicitly illustrates this point.



The second 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_2SO_4 , Na_2CO_3 , $(\text{NH}_4)_2\text{C}_2\text{O}_4$, 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. These observations will be used 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 the elements within this group.

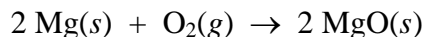
Part 3 – Metal Activity Series

Most elements are classified as metals (see Figure 1). Metals have relatively low ionization energies and tend to lose electrons when forming compounds. Metallic elements often combine with nonmetals to form ionic compounds. For instance, sodium metal (Na) reacts with chlorine gas (Cl_2) to form the ionic compound sodium chloride (NaCl). NaCl is common table salt. Metals also tend to react with oxygen to form metal oxides. For example, iron reacts with oxygen to form rust (Fe_2O_3). 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). As previously discussed, some metals react with water to form hydrogen gas. 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 dis-

placement 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.

Metals reacting with oxygen

Magnesium metal reacts with oxygen to form solid magnesium oxide.



As shown in Figure 5, 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. Oxidation is the loss of one or more electrons. An element that undergoes oxidation shows an increase in oxidation state (*i.e.*, the oxidation state of the element becomes more positive). In this reaction, the oxidation state of magnesium increases from 0 to +2. Oxygen is reduced in this process. Reduction is the gain of electron(s). An element that is reduced undergoes a reduction in its oxidation state. In this case, the oxidation state of oxygen changes from 0 to -2.



Figure 5. Burning magnesium metal in air.

Metals reacting with water

Figure 2c demonstrates what happens when potassium, a Group 1 metal, is placed in 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 6 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.



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

The reactivity of the Group 2 metals towards water increases going down the group, mirroring the trend exhibited by the alkali metals. For instance, the reaction between calcium and water (shown in Figure 6) will occur in cold water; steam is required, however, for the analogous reaction to occur with magnesium. In both cases, the Group 2 metal undergoes oxidation. The species undergoing oxidation in a redox reaction is referred to as the reducing agent. Hence, Ca and Mg are the reducing agents in these processes. The observed differences in water reactivity suggest that calcium is a stronger reducing agent than magnesium and, indeed, this is the case. Strontium reacts with water with even more vigor than magnesium and calcium, and barium is even more reactive towards water than strontium. By combining this information, the four Group 2 metals can be ranked in order of increasing reactivity (or increasing strength as a reducing agent). This ranking would be: Mg

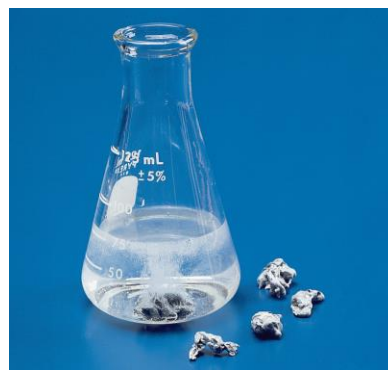
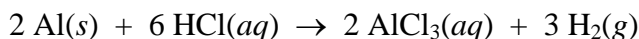


Figure 6. The reaction between calcium and water.

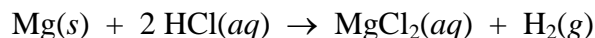
$\text{Ca} < \text{Sr} < \text{Ba}$, indicating that the reactivity of these alkaline earth metals increases going down the group in the periodic table.

Metals reacting with acid

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



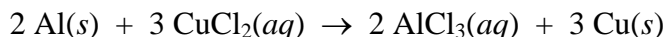
This is another example of a hydrogen displacement redox reaction. In this case, aluminum undergoes oxidation and displaces hydrogen from hydrochloric acid. Many metals are capable of displacing hydrogen from acids. Some metals that react slowly (or not at all) with water react readily with acids to 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 H_2 gas. This reaction is shown below.



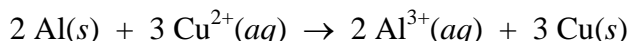
Keep in mind that not all metals react with water or acids to release hydrogen gas. If copper readily reacted with water to form hydrogen gas, household plumbing would certainly not be constructed out of this metal. This example illustrates how the reactivity of a metal strongly influences its application.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

Metal displacement

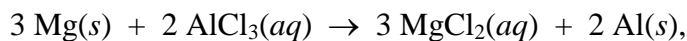
When aluminum metal is added to a solution of copper(II) chloride, the aluminum displaces the copper from the chloride compound to form copper metal.



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.



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:



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). This order was found to be: $\text{Mg} < \text{Ca} < \text{Sr} < \text{Ba}$. 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) 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: $\text{Cu} < \text{Al} < \text{Mg}$. Combining this activity series with the aforementioned Group 2 activity series gives: $\text{Cu} < \text{Al} < \text{Mg} < \text{Ca} < \text{Sr} < \text{Ba}$. Part 3 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

Please read carefully through this lab write-up. Additional information about redox reactions can be found in your textbook. The material presented on pages 382-384 is especially germane to this experiment.

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. 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) *Briefly* describe the objectives of this experiment.
- 2) Write out the experimental procedure in your lab notebook according to the “*Guidelines for Keeping a Laboratory Notebook*” handout.

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

Procedure

Part 1 of this experiment is a “dry-lab” investigation. You might consider doing this exercise last as Parts 2 and 3 of this experiment involve a number of qualitative tests that can take a considerable amount of time if not conducted properly and efficiently.

Part 1 – Periodic Trends in Melting Point Temperatures

Part 1 of this investigation involves exploring periodic trends in the melting points of metals. The data used in this study will be obtained from WebElements, a very informative and creative website that acts as an interactive, online periodic table. Please follow the instructions given below.

1. Open Excel and start a new workbook.
2. Go to the periodicity section of the WebElements website, found at <http://www.webelements.com/periodicity>. This will open up an index that lists a variety of different chemical and physical properties. Scroll down the page and click on the *Melting point* link. Once the new screen updates, you should see an interactive periodic table or chart. This table/chart will allow you to view the melting points of the elements in a variety of different formats. A *shaded table* and *ball chart* are shown by default. Buttons for *cityscape* and *cylinders* renditions are located on the left hand side. As you examine these display options, notice how some representations are more effective at revealing trends than others.
3. In either the *shaded table* or *ball chart* view, you should be able to access the melting point (in K) of an element by simply moving your mouse over the element’s symbol in the periodic table. You can use this method to quickly obtain the melting point temperatures needed in the following steps.
4. *Graph 1* – In your Excel spreadsheet, record the atomic numbers of the Group 1 elements (excluding hydrogen) in one column. Type the melting point temperatures of the corresponding elements in the neighboring column. Use these data to create a “Melting Point vs. Atomic Number” XY scatter plot. Choose the chart sub-type option that graphs data points connected by lines. The melting point temperatures (in K) 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.
5. *Graph 2* – In a second graph, you should plot the melting point temperatures of the Period 3 metals (Na, Mg, and Al) as a function of atomic number. The relevant data should be obtained from Web Elements.
6. *Graphs 3, 4, and 5* – In the third, fourth, and fifth graphs, plot the melting point temperatures of the Period 4 metals (up to Zn), the Period 5 metals (up to Cd), and the Period 6 metals (up to Hg), respectively, as a function of atomic number.

Discussion and Analysis: Discuss the trends seen in Graphs 1-5 and provide a reasonable explanation for these trends based on electron configurations, nuclear charge, shielding, atomic radii, bonding, cohesion energies, and/or crystal packing. Of the metals listed on Graphs 1-5, which has the highest melting point and which has the lowest melting point? Hint: The element with the highest melting point is commonly used in incandescent light bulbs. The element with the lowest melting point is actually a liquid at room temperature and standard atmospheric pressure.

Part 2 – 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 ($\text{C}_2\text{O}_4^{2-}$), 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.

test reagents Group 2 nitrates	H_2SO_4	Na_2CO_3	$(\text{NH}_4)_2\text{C}_2\text{O}_4$	KIO_3
$\text{Ba}(\text{NO}_3)_2$				
$\text{Ca}(\text{NO}_3)_2$				
$\text{Mg}(\text{NO}_3)_2$				
$\text{Sr}(\text{NO}_3)_2$				

2. *Test 1 - Reacting the Group 2 Nitrates with H_2SO_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 H_2SO_4 in each test tube. Add 5 drops of 0.1 M $\text{Ba}(\text{NO}_3)_2$ 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 $\text{Ca}(\text{NO}_3)_2$, 0.1 M $\text{Mg}(\text{NO}_3)_2$, and 0.1 M $\text{Sr}(\text{NO}_3)_2$ 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 Na_2CO_3*

The second test reagent is sodium carbonate. Add 5 drops of 1 M Na_2CO_3 to four different test tubes. Add 5 drops of 0.1 M $\text{Ba}(\text{NO}_3)_2$ 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 $\text{Ca}(\text{NO}_3)_2$, 0.1 M $\text{Mg}(\text{NO}_3)_2$, and 0.1 M $\text{Sr}(\text{NO}_3)_2$ 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 $\text{Ba}(\text{NO}_3)_2$ 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 $\text{Ca}(\text{NO}_3)_2$, 0.1 M $\text{Mg}(\text{NO}_3)_2$, and 0.1 M $\text{Sr}(\text{NO}_3)_2$ 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 KIO_3*

The last test reagent is potassium iodate. Add 5 drops of 0.1 M KIO_3 to four different test tubes. Add 5 drops of 0.1 M $\text{Ba}(\text{NO}_3)_2$ 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 $\text{Ca}(\text{NO}_3)_2$, 0.1 M $\text{Mg}(\text{NO}_3)_2$, and 0.1 M $\text{Sr}(\text{NO}_3)_2$ 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. 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 3 – 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.

1. 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 ta-

ble shown below indicate tests that need not be (and will not be) performed as no displacement reaction will occur.

solns metals	HCl	Ca(NO ₃) ₂	Cu(NO ₃) ₂	FeSO ₄	Fe(NO ₃) ₃	Mg(NO ₃) ₂	SnCl ₄	Zn(NO ₃) ₂
Ca								
Cu								
Fe								
Mg								
Sn								
Zn								

2. Reactions with Hydrochloric Acid

Obtain six small test tubes. Place 10 drops of 6 M HCl in each of the six test tubes. Examine how Ca, Cu, Fe, Mg, Sn, and Zn react with hydrochloric acid by adding a small piece of each metal to a different test tube. Monitor the solutions for evidence of a chemical reaction. For example, look for the evolution of a gas or a persistent color change. Record your observations. If a reaction occurs, the relative speed of the reaction should be noted (*i.e.*, indicate whether the reaction was immediate, slow, very slow, etc.). If a gas is produced, indicate whether the reaction was vigorous or weak.

3. 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 than 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. In your discussion, you should explain in detail how you arrived at your final prediction regarding the relative reactivities of these metals. In other words, you must indicate how you came up with your metal activity series.

Report

Your lab report should be a formal, individual report that addresses all the questions posed in this write-up, including any listed in the *Discussion and Analysis* sections. This document should be prepared according to the instructions given in the “Guidelines for Laboratory Reports” handout. In addition to the categories discussed in these guidelines you should provide answers to all the questions posed in this laboratory experiment write-up.

References

This experiment is modified from the following references:

“Experiment 12 The Alkaline Earths and Halogens – Two Families in the Periodic Table”, in *Chemical Principles in the Laboratory 9th Edition*, E. J. Slowinski, W. C. Wolsey, Brooks/Cole, Belmont CA (2009).

“Experiment 15 Activity Series” in *Chemistry the Central Science Laboratory Experiments*, J. H. Nelson, K. C. Kemp, Prentice Hall, Upper Saddle River, NJ (2000).