

1. Properties of Some Representative Elements

Objective

In this experiment you will examine the properties of some of the more common representative elements.

Introduction

Even in the earliest studies of chemistry, it became evident that certain elemental substances were very much like other substances in their physical and chemical properties. For example, the more common alkali metals (Na, K) were almost indistinguishable to early chemists. Both metals are similar in appearance and undergo reaction with the same reagents (giving formulas of the same stoichiometric ratio). As more and more elemental substances were separated, purified, and identified, more and more similarities between the new elements and previously known elements were detected. Chemists began to wonder why such similarities existed.

In the mid-1800s, Mendeleev and Meyer independently proposed that the properties of the known elements seemed to vary systematically when arranged in the order of their atomic masses. That is, when a list of the elements is made in order of increasing atomic mass, a given set of properties would repeat at regular intervals among the elements. Mendeleev generally receives most credit for the development of the periodic law, because he suggested that some elements were missing from the table. That is, based on the idea that the properties of the elements should repeat at regular intervals, Mendeleev suggested that some elements that should have certain properties had not yet been discovered. Mendeleev went so far as to predict the properties of the as yet undiscovered elements, based on the properties of the already known elements. Mendeleev's periodic law received much acceptance when the missing elements were finally discovered and had the very properties that Mendeleev had predicted for them.

While Mendeleev arranged the elements in order of increasing atomic mass, the modern periodic table is arranged in order of increasing atomic number. At the time of Mendeleev's work, the structure of the atom had not yet been determined. Generally, the arrangement of the elements by atomic mass is very similar to the arrangement by atomic number, with some notable exceptions. Some elements in Mendeleev's arrangement were out of order and did not have the properties expected. When the arrangement is made by atomic number, however, the properties of the elements do fall into a completely regular order.

Group 1A Elements

The Group 1A elements are commonly referred to as the **alkali metals**. These substances are among the most reactive of elements, are never found in nature in the uncombined state, and are relatively difficult and expensive to produce and store. In particular, these elements are very easily oxidized by oxygen in

the air; usually they are covered with a layer of oxide coating unless the metal has been freshly cut or cleaned. When obtained from a chemical supply house, these elements are usually stored under a layer of kerosene to keep them from contact with the air.

The elements of Group 1A are very low in density and are soft enough to be easily cut with a knife or spatula. The only real reaction undergone by these elements involves the loss of their single valence electron to some other species:



In compounds, the elements of Group 1A are invariably found as the unipositive ion. The least reactive of these elements is lithium (the topmost member, with the valence electron held most tightly by the nuclear charge), whereas the most reactive is cesium (the bottommost member). Cesium is so reactive that the metal may be ionized merely by shining light on it. This property of cesium has been made use of in certain types of photocells (in which the electrons produced by the ionization of cesium are channeled into a wire as electric current). The most common elements of this group, sodium and potassium, are found naturally in great abundance in the combined state on the earth. The elements of Group 1A are rather low in density (for example, they float on the surface of water while reacting) and are very soft (they can be easily cut with a knife).

Group 2A Elements

The elements of Group 2A are commonly referred to as the **alkaline earth elements**. The “earth” in this sense indicates that the elements are not as reactive as the alkali metals of Group 1A and are found (occasionally) in the free state in nature. Beryllium, the topmost element of this group, is not very common. But some of the other members of the group, such as magnesium and calcium, are found in fairly high abundance. As with the Group 1A elements, the relative reactivity of the Group 2A elements increases from top to bottom in the periodic table. For example, metallic magnesium can be kept almost indefinitely without developing an oxide coating (by reaction with air), whereas a freshly prepared sample of pure metallic calcium will develop a layer of oxide in a much shorter period. In compounds, the elements of Group 2A are usually found as dipositive ions (through loss of the two valence electrons).

Group 3A Elements

Whereas all the elements of Group 1A and Group 2A are metals, the first member of Group 3A (boron) is a nonmetal. The dividing line between metallic elements and nonmetallic elements forms the “stairstep” region indicated at the right-hand side of most periodic charts. Because boron is to the right of the stairstep, boron shows many properties that are characteristic of nonmetals. For example, boron is not a very good conductor of heat and electricity, whereas aluminum (which is the element beneath boron in the group) shows the much higher conductivities associated with metallic elements. Boron is generally found covalently bonded in its compounds, which further indicates its nonmetallic nature. Aluminum is the only relatively common element from Group 3A.

In nature, aluminum is generally found as the oxide, and most metallic aluminum is produced by electrolysis of the molten oxide. This is a relatively difficult and expensive process, which is part of the reason the aluminum in cans was commonly recycled even before the present emphasis on preserving the environment. Aluminum is a relatively reactive element; however, this reactivity is sometimes masked under ordinary conditions. Aluminum used for making cooking utensils or cans becomes very quickly coated with a thin layer of aluminum oxide. This oxide layer serves as a protective coating, preventing further oxidation of the aluminum metal underneath. (This is why aluminum pans are seldom as shiny as copper or stainless steel pans.)

Group 4A Elements

Group 4A contains some of the most common, useful, and important elements known. The first member of the group, carbon, is a nonmetal. The second and third members, silicon and germanium, show both some metallic and some nonmetallic properties (and because of this in-between nature, these elements are referred to as **metalloids** or **semimetals**). The last two members of the group, tin and lead, show mostly metallic properties.

The element carbon forms the framework for life. Virtually all biological molecules are, in fact, carbon compounds. Carbon is unique among the elements in that it is able to form long chains of many hundreds or even thousands of similar carbon atoms. Pure elemental carbon itself is usually obtained in either of two forms, graphite or diamond, each of which has a very different structure. Graphite contains flat, two-dimensional layers of covalently bonded carbon atoms, with each layer effectively being a molecule of graphite, independent of the other layers present in the sample. The layers of carbon atoms in graphite are able to move relative to each other. This makes graphite slippery and useful as a lubricant for machinery and locks. Diamond contains three-dimensional extended covalent networks of carbon atoms, making an entire diamond effectively a single molecule. Because all the atoms present in diamond are covalently bonded to each other in three dimensions, diamond is relatively stable and is also the hardest substance known. Natural diamonds used as gemstones are produced deep in the earth under high pressures over eons, but synthetic diamonds are produced in the laboratory by compressing graphite to several thousand atmospheres of pressure. Because of its hardness, diamond is used as an abrasive in industry.

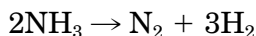
Silicon and germanium are used in the semiconductor industry. Because these elements have some of the properties of nonmetals (brittleness, hardness), as well as some of the properties of metals (slight electrical conductivity), they have been used extensively in transistors.

The last two elements of Group 4A, tin and lead, show mostly metallic properties (shiny luster, conductivity, malleability) and frequently form ionic compounds. However, covalently bonded compounds of these metals are not rare. For example, leaded gasoline contains tetraethyllead, which is effectively an organic lead compound. Tin has been used in the past to make cooking utensils and is used extensively as a liner for steel cans ("tin" cans), because it is less likely to be oxidized than iron and is also less likely to impart a bad taste to foods. Whereas tin is nonpoisonous and is used in the food industry, lead compounds are very dangerous poisons. Formerly, pigments for white paint

were made using lead oxide. Lead compounds are remarkably sweet-tasting, and many children have been seriously poisoned by eating flecks of peeling lead paints.

Group 5A Elements

The first few elements of Group 5A are nonmetals, with nitrogen and phosphorus being the two most abundant elements of the group. The earth's atmosphere is nearly 80% N₂ gas by volume. This large percentage of nitrogen is believed to be a remnant of the primordial composition of the earth's atmosphere. Astronomers have determined that several of the planets of the solar system contain large quantities of ammonia, NH₃, in their atmospheres. It is known that ammonia actually comes to equilibrium with its constituent elements



Astronomers theorize that the atmosphere of the early earth was also mostly ammonia, but that the lighter hydrogen gas has escaped the earth's relatively weak gravitational field over the eons, leaving the atmosphere with a large concentration of nitrogen gas. Nitrogen is also an important constituent of many of the molecules synthesized and used by living cells (amino acids, for example).

Phosphorus, the second member of Group 5A, is a typical solid nonmetal. Three allotropic forms of elemental phosphorus exist: red, white, and black phosphorus. White phosphorus is the most interesting of the three forms and is the form that is produced when phosphorus vapor is condensed. White phosphorus is very unstable toward oxidation. Below about 35°C, white phosphorus reacts with oxygen of the air, emitting the energy of the oxidation as light rather than heat (phosphorescence); above 35°C, white phosphorus will spontaneously burst into flame. (For this reason, white phosphorus is stored under water to keep it from contact with air.)

Whereas elemental nitrogen consists of diatomic molecules, white phosphorus contains tetrahedrally shaped P₄ molecules. When white phosphorus is heated to high temperatures in the absence of air, it is converted to the red allotrope. The red allotrope is much less subject to oxidation and is the form usually available in the laboratory. The lower reactivity is due to the fact that the phosphorus atoms have polymerized into one large molecule, which is far less subject to attack by oxygen atoms than the individual P₄ molecules of the white allotrope.

Phosphorus is essential to life, being a component of many biological molecules, most notably the phosphate compounds used by the cell for storing and transferring energy (AMP, ADP, and ATP). Phosphorus also forms part of the structural framework of the body, bones being complex calcium phosphate compounds.

The other members of Group 5A are far less common than are nitrogen and phosphorus. For example, arsenic (As) is known primarily for its severe toxicity and has been used in various rodenticides and weed killers (though As has been supplanted by other agents for most of these purposes).

Group 6A Elements

The most abundant elements of Group 6A (sometimes referred to as the **chalcogens**) are oxygen and sulfur, which, along with selenium, are non-metallic. Elemental oxygen makes up about 20% of the atmosphere by volume and is vital not only to living creatures but to many common chemical reactions. Most of the oxygen in the atmosphere is produced by green plants, especially by plankton in the oceans of the earth. On a microscopic basis, oxygen is used in living cells for the oxidation of carbohydrates and other nutrients. Oxygen is needed also for the oxidation of petroleum-based fuels for heating and lighting purposes and for myriad other important industrial processes. Oxygen has two allotropic forms: the normal elemental form of oxygen is O₂ (dioxygen), whereas a less stable allotrope, O₃ (ozone), is produced in the atmosphere by high-energy electrical discharges.

Sulfur is obtained from the earth in a nearly pure state and needs little or no refining before use. In certain areas of the earth, there are vast underground deposits of pure sulfur. High-pressure steam is pumped into such deposits, liquefying the sulfur and allowing it to be pumped to the surface of the earth for collection. The main use of sulfur is in the manufacture of sulfuric acid, which is the industrial chemical produced each year in the largest amount. When sulfur is burned in air, the sulfur is converted to sulfur dioxide, notable for its choking, irritating odor (similar to a freshly struck match). If sulfur dioxide is further oxidized with a catalyst, it is converted to sulfur trioxide, the anhydride of sulfuric acid.

Sulfur has several allotropic forms. At room temperature, the normal allotrope is orthorhombic sulfur, which consists of S₈ cyclic molecules. At higher temperatures, a slight rearrangement of the shape of these rings occurs, producing a solid allotropic form of sulfur: monoclinic sulfur. This substance has a different crystalline structure than the orthorhombic form. (Monoclinic sulfur slowly changes into orthorhombic sulfur at room temperature.) The most interesting allotrope of sulfur, however, is produced when boiling sulfur is rapidly lowered in temperature. This technique produces a plastic (amorphous) allotrope. The properties of plastic sulfur are very different from those of orthorhombic or monoclinic sulfur. Whereas plastic sulfur is soft and pliable (being able to be stretched and pulled into strings), orthorhombic and monoclinic sulfur are both hard, brittle solids. Sulfur is insoluble in water, but the orthorhombic form is fairly soluble in the nonpolar solvent carbon disulfide, CS₂.

The most common oxygen compound is, of course, water, H₂O. The analogous sulfur compound, hydrogen sulfide, H₂S, is a noxious gas at room temperature. (The gas smells like rotten eggs; eggs contain a protein involving sulfur that releases hydrogen sulfide as the egg spoils.)

Selenium, which occurs in Group 6A just beneath sulfur, is the only other member of the group that occurs in any abundance. Selenium compounds are generally very toxic. Selenium is used in certain prescription dandruff shampoos.

Group 7A Elements

The Group 7A elements are more commonly referred to as the **halogens** (“salt-formers”). The elementary substances are very reactive; these elements are, therefore, usually found in the combined state, as the negative halide ions. Elemental fluorine is the most reactive nonmetal and is seldom encountered in the laboratory because of problems of toxicity and handling of the gaseous element. Most compounds of fluorine are very toxic, although tin(II) fluoride is used in toothpastes and mouthwashes as a decay preventative.

Hydrogen fluoride is very different from the other hydrogen halides, since aqueous solutions of HF are weakly acidic, rather than strongly acidic (HCl, HBr). Hydrogen fluoride is able to attack and dissolve glass and is always stored in plastic bottles.

Gaseous elemental chlorine is somewhat less reactive than elemental fluorine and is used commercially as a disinfectant. Chlorine will oxidize biological molecules present in bacteria (thereby killing the bacteria), while it is reduced itself to chloride ion (which is nontoxic). For example, drinking water is usually chlorinated, either with gaseous elemental chlorine or with some compound of chlorine that is capable of releasing elemental chlorine when needed. Gaseous hydrogen chloride, when dissolved in water, forms the solution known as hydrochloric acid, which is one of the most commonly used acids.

Elemental bromine is a strikingly dark red liquid at room temperature. Elemental bromine is used as a common laboratory test for the presence of double bonds in carbon compounds. Br₂ reacts with the double bond, and its red color disappears, thereby serving as an indicator of reaction.

Elemental iodine is a dark gray solid at room temperature but undergoes sublimation when heated slightly, producing an intensely purple vapor. Elemental iodine is the least reactive of the halogens. Iodine is very important in human metabolism, being a component of various hormones produced by the thyroid gland. Since many people’s diets do not contain sufficient iodine from natural sources, commercial table salt usually has a small amount of sodium iodide added as a supplement (“iodized” salt). Iodine is also used sometimes as a topical antiseptic, since it is a weak oxidizing agent that is capable of destroying bacteria. Iodine for this purpose (called “tincture” of iodine) is usually sold in alcohol solution.

SAFETY PRECAUTIONS

- **Wear safety glasses at all times while in the laboratory.**
- **The reactions of lithium, sodium, and potassium with water are very dangerous. Your instructor will demonstrate these reactions for you. Do *not* attempt these reactions yourself.**
- **Hydrogen gas is extremely flammable and forms explosive mixtures with air. Small amounts of hydrogen will be generated in some of the student reactions. Avoid flames and use caution.**

- **The flames produced by burning magnesium/calcium are intensely bright and can damage the eyes. Do *not* look directly at the flame while these substances are burning.**
- **Oxides of nitrogen are toxic and extremely irritating to the respiratory system. Generate these gases only in the exhaust hood.**
- **Hydrochloric and nitric acids will burn the skin. Wash immediately if the acids are spilled and inform the instructor.**
- **Vapors of methylene chloride and of the halogens are toxic. Keep these substances in the exhaust hood during use. Dispose of methylene chloride as directed by the instructor.**
- **Hydrogen sulfide and sulfur dioxide are toxic and have noxious odors. Confine them to the exhaust hood.**
- **Salts of the Group 1A and 2A metals may be toxic; wash after handling these compounds.**

Apparatus/Reagents Required

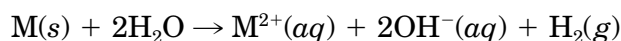
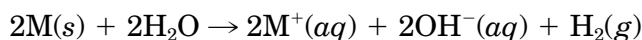
Lithium, sodium, potassium, universal indicator solution, calcium, magnesium ribbon, samples of solid LiCl, NaCl, KCl, CaCl₂, BaCl₂, and SrCl₂, flame test wires, 6 M hydrochloric acid, boric acid, aluminum oxide, sodium carbonate, copper wire, concentrated nitric acid, 0.1 M NaOH, chlorine water, bromine water, iodine water, methylene chloride, 0.1 M NaCl, 0.1 M NaBr, 0.1 M NaI, ferrous sulfide, solid iodine

Procedure

Record all data and observations directly in your notebook in ink.

A. *Some Properties of the Alkali and Alkaline Earth Elements*

The reactive metals of Group 1A and Group 2A react with cold water to liberate elemental hydrogen, leaving a solution of the strongly basic metal hydroxide



Generally the reactivity of the metals increases going from the top of the group toward the bottom of the group. The valence electrons of the metal are less tightly held by the nucleus toward the bottom of the group.

1. *Lithium, Sodium, and Potassium (Instructor Demonstration)*

The reaction of the Group 1A metals with water is very dangerous and will be performed by your instructor as demonstrations.

In the exhaust hood, behind a safety shield, the instructor will drop small pellets of lithium, sodium, and potassium into beakers containing a small amount of cold water. Compare the speed and vigor of the reactions.

When the reactions have subsided, obtain small portions of the water from the beakers in which the reactions were conducted in separate clean test tubes. Add 2–3 drops of universal indicator to each test tube; refer to the color chart provided with the indicator to determine the pH of the solution. Write equations for the reactions of the metals with water.

2. *Reactions of Magnesium and Calcium (Student Procedure)*

Repeat the procedure that you have seen in the demonstration, using pieces of the Group 2A metals magnesium and calcium in place of lithium/sodium/potassium. Work in the exhaust hood, with the safety shield pulled down. Account for differences in reactivity between these two Group 2A metals. Test the pH of the water with which the metals were reacted.

Add a single turning of magnesium to 10 drops of 6 M HCl in a small test tube. Why does magnesium react with *acid* but not with water?

3. *Flame Tests*

Obtain a 6–8-inch length of nichrome wire for use as a flame test wire. Form one end of the wire into a small loop that is no more than a few millimeters in diameter.

Obtain about 20 mL of 6 M HCl in a small beaker, and immerse the loop end of the wire for 2–3 minutes to clean the wire.

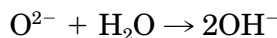
Ignite a Bunsen burner flame, and heat the loop end of the wire in the flame until it no longer imparts a color to the flame.

Obtain a few crystals each of LiCl, NaCl, KCl, CaCl₂, BaCl₂, and SrCl₂. Dip the loop of the flame test wire into one of the salts, and then into the oxidizing portion of the burner flame. Record the color imparted to the flame by the salt.

Clean the wire in 6 M HCl, heat in the flame until no color is imparted, and repeat the flame test with the other salts. The colors imparted to the flame by the metal ions are so intense and so characteristic that they are frequently used as a test for the presence of these elements in a sample.

B. Oxides/Sulfides of Some Elements

Oxygen compounds of most elements are known, and are generally referred to as **oxides**. However, there are great differences between the oxides of metallic elements and those of nonmetals. The oxides of metallic elements are *ionic* in nature: they contain the oxide ion, O²⁻. Ionic metallic oxides form *basic* solutions when dissolved in water



Oxides of nonmetals are generally *covalently* bonded and form *acidic* solutions when dissolved in water.

Sulfur compounds of both metallic and nonmetallic substances are known. Sulfur compounds of nonmetallic compounds are generally covalently bonded, whereas with metallic ions, sulfur is usually present as the sulfide ion, S^{2-} . Most sulfur compounds have characteristic unpleasant odors, or decompose into compounds that have the characteristic odor.

1. *Metallic Oxides*

Place a small amount of distilled water in a beaker and add 2–3 drops of universal indicator. Refer to the color chart provided with the indicator, and record the pH of the water.

Obtain a piece of magnesium ribbon about 1 inch in length. Hold the magnesium ribbon with tongs above the beaker of water and ignite the metal with a burner flame. Do *not* look directly at the flame produced by burning magnesium. It is intensely bright and *damaging* to the eyes. When the flame has expired, stir the liquid in the beaker for several minutes and record the pH of the solution. Write an equation for the reactions that have taken place.

Place approximately 10 drops of distilled water in each of three test tubes, and add a drop of universal indicator to each. Add a *very* small quantity of sodium (per)oxide to one test tube, calcium oxide (lime) to a second test tube, and aluminum oxide (alumina) to the third test tube. Stir each test tube with a clean glass rod, record the pH of the solution, and write an equation to account for the pH.

2. *Nonmetallic Oxides*

Place approximately 10 drops of water into each of two test tubes, and add 1 drop of universal indicator to each. In one test tube, dissolve a small amount of boric acid. To the second test tube, add a small chunk of dry ice (solid carbon dioxide). Stir the test tubes with a clean glass rod, record the pH, and write an equation to account for the pH of the solutions.

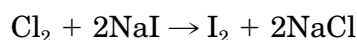
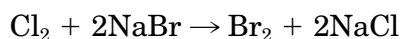
3. *Sulfur Compounds*

Have ready a beaker of cold water. In the exhaust hood, ignite a *tiny* amount of powdered sulfur on the tip of a spatula in the burner flame and *cautiously* note the odor. Extinguish the sulfur in the cold water to prevent too much SO_2 from getting into the room air. Determine the pH of the water in which you extinguished the sulfur. Why is it acidic?

Obtain a *tiny* portion of iron(II) sulfide in a test tube, and add 2 drops of dilute hydrochloric acid. *Cautiously* note the odor of the hydrogen sulfide generated. Transfer the test tube to the exhaust hood to dispose of the hydrogen sulfide.

C. *Some Properties of the Halogen Family (Group 7)*

All of the members of Group 7A are nonmetallic. The halogen elements tend to gain electrons in their reactions, and the attraction for electrons is stronger if the electrons are closer to the nucleus of the atom. The activity of the halogen elements therefore decreases from top to bottom in Group 7 of the periodic table. As a result of this, elemental chlorine is able to replace bromide and iodide ion from compounds:



Similarly, elemental bromine is able to replace iodide ion from compounds.

1. *Identification of the Elemental Halogens by Color*

The elemental halogens are nonpolar and are not very soluble in water. Solutions of the halogens in water are not very brightly colored. However, if an aqueous solution of an elemental halogen is shaken with a nonpolar solvent, the halogen is preferentially extracted into the nonpolar solvent, and imparts a characteristic, relatively bright color to the nonpolar solvent. Since the nonpolar solvent is not miscible with water, the halogen color is evident as a separate colored layer.

In the following tests, dispose of the methylene chloride samples as directed by the instructor.

Obtain about 1 mL of chlorine water in a small test tube. Note the color. In the exhaust hood, add 10 drops of methylene chloride to the test tube. Stopper and shake.

Allow the solvent layers to separate and note the color of the lower (methylene chloride) layer. Elemental chlorine is not very intensely colored and imparts only a pale yellow/green color.

Repeat the procedure using 1 mL of bromine water in place of the chlorine water. Bromine imparts a bright red color to the lower layer.

Repeat procedure using 1 mL of iodine water. Iodine imparts a purple color to the methylene chloride layer.

2. *Relative Reactivity of the Halogens*

Add 2 mL of chlorine water to each of two test tubes. Add 2 mL of 0.1 M NaBr to one test tube and 2 mL of 0.1 M NaI to the second tube.

In the exhaust hood, add 10 drops of methylene chloride to each test tube. Stopper, shake, and allow the solvents to separate. Record the colors of the lower layers and identify which elemental halogens have been produced.

Add 2 mL of bromine water to each of two test tubes. Add 2 mL of 0.1 M NaI to one test tube and 2 mL of 0.1 M NaCl to the other test tube.

In the exhaust hood, add 10 drops of methylene chloride to each test tube. Stopper, shake, and allow the layers to separate. Record the colors of the lower layers. Indicate where a reaction has taken place.

Add 2 mL of iodine water to each of two test tubes. Add 2 mL of 0.1 M NaBr to one test tube and 2 mL of 0.1 M NaCl to the other test tube.

In the exhaust hood, add 10 drops of methylene chloride to each test tube. Stopper, shake, and allow the layers to separate. Record the colors of the lower layers. Did any reaction take place?

3. *Sublimation of Iodine*

Using forceps, place a few crystals of elemental iodine in a small beaker. Set the beaker on a ring stand in the exhaust hood, and cover with a watch glass containing some ice. Heat the iodine crystals in the beaker with a small flame for 2–3 minutes, and watch the sublimation of the iodine. Extinguish the flame, and examine the iodine crystals that have formed on the lower surface of the watch glass.

Properties of Some Representative Elements

Date: Student name:
Course: Team members:
Section:
Instructor:

Prelaboratory Questions

1. For each of the following elements, list the atomic number, average atomic mass, in which group (vertical column) of the periodic table the element can be found, and also in which period (horizontal row) the element is located:

K

Ba

I

Ne

Se

Cs

Ra

Kr

2. Why do members of the same vertical group of the periodic chart tend to show similar chemical and physical properties?

Properties of Some Representative Elements

Date: Student name:
Course: Team members:
Section:
Instructor:

Results/Observations

A. Some Properties of the Alkali and Alkaline Earth Elements

1. Reaction with water

Metal	Evidence of reaction	Relative vigor	Indicator color	Equation for reaction
Li
Na
K
Mg
Ca

2. Reaction of Mg with acid.....

3. Flame tests

Metal	Color observed
.....
.....
.....
.....
.....
.....

B. Oxides/Sulfides of Some Elements

Substance	Results/equations—universal indicator test
Magnesium oxide
Sodium oxide
Calcium oxide
Aluminum oxide
Boric acid
Dry ice
Sulfur dioxide
FeS + HCl

C. Some Properties of the Halogen Family

1. Identification of the elemental halogens

Halogen	Color of water solution	Color of CH ₂ Cl ₂ extract
Cl ₂
Br ₂
I ₂

2. Relative reactivity of the halogens

Elemental halogen	Halide ion	Color of lower layer	Which halogen is in lower layer?	Did a reaction take place?
Cl ₂	Br ⁻
Cl ₂	I ⁻
Br ₂	I ⁻
Br ₂	Cl ⁻
I ₂	Br ⁻
I ₂	Cl ⁻

Student name: Course/Section: Date:

Equations for reactions that occurred

Sublimation of iodine observation.....

Questions

1. In comparing the *vigor* of the reactions of lithium, sodium, and potassium with cold water, does there appear to be a correlation between the location of these elements in the periodic table and the vigor of reaction?
2. You tested solutions of several oxides with universal indicator to determine their acidity/basicity. Can a *general* statement be made concerning the acidity of metal oxides compared to oxides of nonmetallic elements?
3. Elemental fluorine, F_2 , was not used in the halogens experiment because of difficulties in handling it, and elemental astatine, At_2 , was not used because of its radioactivity and short half-life. How would you expect the reactivity of F_2 and At_2 to compare with those of the elemental halogens tested in this experiment?

