General Chemistry Lab
Experiment 2

Reactivity of Metals

Introduction

About three-fourths of the 112 elements are metals. Of these, at least 20 are fairly common and are of major importance. To get acquainted with some physical properties of various metals peruse table 1. You will study some of the chemical properties of metals in today's experiment. Metals in their elemental form are commonly described as lustrous. They are good conductors of heat and electricity and are capable of being deformed without breaking. The fusing (melting together) of particular metals gives alloys which have a wide variety of properties and uses. Steels are alloys of iron with other metals and carbon. Stainless steels contain nickel and chromium. Brass contains mostly copper and zinc, whereas bronze is primarily copper and tin.

Metals and nonmetals combine in many ways to make compounds, some of which are soluble in water. In most compounds the metal exists in the form of positive ions and the nonmetal as negative ions. Generally those compounds of the metals whose negative ions are oxides, $O^{2-}$, or hydroxides, $OH^-$ are called bases. The others are called salts. Examples, showing the charges on the ions, are sodium chloride, $Na^+ Cl^-$, and copper sulfate, $Cu^{2+} SO_4^{2-}$. Keep in mind the distinction between the free metal (the element) and the combined form of the metal (the compound); the former consists of neutral metal atoms and the latter contains positively charged metal ions (along with negative nonmetal ions). There is a vast difference between copper metal (red, lustrous, solid, insoluble in water, somewhat soft and flexible) and copper ions in copper sulfate (blue crystal, brittle, soluble in water). When you use copper sulfate solution in this experiment, note how copper (as ions) in solution is very unlike copper (as atoms) in metallic form. This experiment focuses on the properties of free metals.

In nature metals usually are found in combination with oxygen or other nonmetals. Only a few of the less reactive ones, such as gold and sometimes silver and copper, are found in the free metallic (elemental) state.

One of the most important characteristics of a metal is its activity (reactivity, or ability to react to form compounds). Metals range widely in activity, from vigorously reactive cesium, potassium, and sodium, to quite inactive (inert) platinum, gold, and silver. Since the latter resist oxidation, they are often called the noble metals. The coinage metals include gold and silver along with some metals of lesser value, such as copper and nickel. Many metals become oxidized by reacting with oxygen in the air to form a tarnish or rust (oxide). Examples of metallic oxides are sodium oxide, $Na_2O$, aluminum oxide, $Al_2O_3$, ferric oxide, $Fe_2O_3$, and cupric oxide, $CuO$. 
TABLE 1  Some Metals and Their Properties

<table>
<thead>
<tr>
<th>Metal</th>
<th>Symbol</th>
<th>Property</th>
<th>Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Iron</td>
<td>Fe</td>
<td>Strong, tough, corrodes</td>
<td>Main ingredient in steel</td>
</tr>
<tr>
<td>Nickel</td>
<td>Ni</td>
<td>Resists corrosion</td>
<td>Coinage, alloy for stainless steel</td>
</tr>
<tr>
<td>Chromium</td>
<td>Cr</td>
<td>Shiny, resists corrosion</td>
<td>Chrome plating</td>
</tr>
<tr>
<td>Zinc</td>
<td>Zn</td>
<td>Forms protective coating of ZnO</td>
<td>Galvanizing coating</td>
</tr>
<tr>
<td>Radium</td>
<td>Ra</td>
<td>Radioactive</td>
<td>Cancer treatment</td>
</tr>
<tr>
<td>Aluminum</td>
<td>Al</td>
<td>Light and strong</td>
<td>Airplanes, window frames</td>
</tr>
<tr>
<td>Uranium</td>
<td>U</td>
<td>Fissionable</td>
<td>Energy source</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg</td>
<td>Light and strong</td>
<td>Auto wheels, luggage</td>
</tr>
<tr>
<td>Mercury</td>
<td>Hg</td>
<td>Dense liquid at room temperature</td>
<td>Thermometers, barometers</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>Very active, soft</td>
<td>Heat transfer medium</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu</td>
<td>Red color, good electrical conductor</td>
<td>Electrical wiring</td>
</tr>
<tr>
<td>Silver</td>
<td>Ag</td>
<td>Excellent electrical conductor</td>
<td>Electrical contacts, mirrors, jewelry</td>
</tr>
<tr>
<td>Gold</td>
<td>Au</td>
<td>Yellow metal, soft</td>
<td>Coinage, instruments, dentures</td>
</tr>
<tr>
<td>Platinum</td>
<td>Pt</td>
<td>Inert, high melting</td>
<td>Jewelry, instruments</td>
</tr>
<tr>
<td>Tantalum</td>
<td>Ta</td>
<td>Resists attack by acids</td>
<td>Synthetic skulls and bone parts</td>
</tr>
<tr>
<td>Tungsten</td>
<td>W</td>
<td>Very high melting</td>
<td>Light-bulb filament</td>
</tr>
<tr>
<td>Tin</td>
<td>Sn</td>
<td>Resists corrosion</td>
<td>Coating for steel cans</td>
</tr>
<tr>
<td>Lead</td>
<td>Pb</td>
<td>Low melting, dense, soft</td>
<td>Plumbing, fishing weights</td>
</tr>
</tbody>
</table>

Replacement Series

In this experiment you will be ranking some of the metals according to their activities, from most to least active. One way to do this is to observe the relative vigor of their reactions (tendency to react) with nonmetals such as oxygen or chlorine. Another method, which differentiates more clearly, is to note whether one metal can replace another in a chemical compound. The general rule is that the more active metal replaces the less active one. This is easily observed in reactions between metals and metal ions (cations) in solution. Iron, for example, replaces copper from cupric chloride, CuO₂, because it is more active than copper.

\[
Fe + CuCl_2 \rightarrow FeCl_2 + Cu
\]

In this example two electrons are transferred from the neutral iron atom to the positive copper ion. This is more apparent if the equation is written in net ionic form.

\[
Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu
\]

This reaction, which involves the transfer of electrons, can be divided into two half-reactions which show exactly where electrons are gained and lost.

\[
Fe \rightarrow Fe^{2+} + 2e^- \quad \text{and} \quad Cu^{2+} + 2e^- \rightarrow Cu
\]
The general reaction for the replacement of a metal ion by another metal may be written as follows. (We are arbitrarily showing only one electron per atom being transferred.)

\[
\begin{align*}
M_A^- & + M_B^+ \rightarrow M_A^+ + M_B^- \\
\text{Free metal A} & \quad \text{Salt of metal B} & \quad \text{Salt of metal A} & \quad \text{Free metal B}
\end{align*}
\]

Or, in net ionic form,

\[
\begin{align*}
M_A^- & + M_B^+ \rightarrow M_A^+ + M_B^-
\end{align*}
\]

**Replacement of Hydrogen from Acids and Water**

Acids are an important group of compounds, which, in water, produce hydronium ions, \(H_3O^+\). Hydronium ions enter into replacement reactions with metals; so hydrogen appropriately is included in an activity or replacement ranking among metals. Those metals more active than hydrogen will replace hydrogen from acids, but those less active are unable to do so.

\[
\begin{align*}
\text{Zn} + 2\text{HCl} & \rightarrow \text{ZnCl}_2, + \text{H}_2(g) \\
\text{Or, in ionic style,} & \\
\text{Zn} + 2 \text{H}^+ & \rightarrow \text{Zn}^{2+} + \text{H}_2(g)
\end{align*}
\]

You may consider water to be a special case of a very weak acid. Metals which replace hydrogen from acids also liberate hydrogen from water, but drastic conditions may be required. Only the most active metals, such as sodium and potassium, replace hydrogen from water readily at room temperature. The reaction of sodium with water is one you will observe today. This vigorous reaction can be sufficiently exothermic to ignite the explosive hydrogen gas, so you will use only a small amount of sodium. Some reactions of metals with water are

\[
\begin{align*}
2\text{Na} & + 2\text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2 @ 25^\circ\text{C} \\
3 \text{Fe} & + 4 \text{H}_2\text{O} \rightarrow \text{Fe}_3\text{O}_4 + 8\text{H}_2 @ 800^\circ\text{C}
\end{align*}
\]

**Metal Activities and the Periodic Chart**

The activity of a metal is related to the ease with which its outer electrons can be removed. Generally, atoms having the fewest electrons in the outer shell and those with the largest radii lose their electrons most easily. In the periodic chart, elements with the fewest outer electrons are located toward the left side, and those with the largest radii are found toward the bottom. In general, the most active metals are located toward the lower left corner of the chart.

In this experiment you will become familiar with the properties of various metals, and especially with the differences in their reactivities, learn to relate the activity of a metal to its position in the periodic chart and understand clearly the distinction between free metals and metals in the combined state.
**Procedure**

1. **Reaction of Metals with Oxygen**

In this part of the experiment you will make a variety of observations on the reactivity of oxygen in the air with several metals. Record your observations in the table and answer the questions.

   A. **Reaction at Room Temperature**

   DEMONSTRATION: Your instructor will place small strips of metals on a sheet of paper and label them. Each metal will be scratched to expose a clean surface to the air. Suggested metals are iron, tin, lead, copper, magnesium, aluminum, and zinc. Observe the metals throughout the laboratory period. Do they react with oxygen, that is, do they become tarnished?

   Your instructor will cut a piece of sodium. Notice how easily sodium can be cut with a knife. Observe a freshly cut piece and see what happens when air comes into contact with the metal. Does a reaction occur between sodium and oxygen in the air?

   B. **Reaction at Elevated Temperature**

   Most of your metal samples, which you placed on the paper towel, probably have not reacted noticeably with air. Reactions that are sluggish at room temperature often can be accelerated by raising the temperature. Collect another set of metal strips used in #1. Heat the samples by cautiously holding them in the flame with tongs. Do any of the metals burn (react) with oxygen? Could you suggest an ingredient for making flares?

   C. **Effect of Surface Area**

   To determine whether the amount of surface area has an effect on the rate of reaction, heat a bulky piece of iron, such as a nail, and also a sample of iron with a large surface area, such as steel wool. Which of the two reacts most readily with oxygen? After burning in air, the steel wool may still hold its shape. Take some between your fingers when it has **cooled**. Is it still metallic iron?

2. **Reaction of Metals with Water**

   To determine whether the metals of part 1 react with water, place a piece of each, the size of a match head, in 1 mL of water in a test tube. Remember that water acts as an extremely weak acid and that the evidence for a reaction is the evolution of hydrogen gas (bubble formation). DEMONSTRATION: Your instructor will cautiously add a small piece of sodium to water in a beaker. Record in the table your observations for sodium and the other metals and answer the questions.

3. **Reaction of Metals with an Acid**

   Now repeat the reaction of part 2 with three different concentrations of a strong acid, such as hydrochloric acid, HCl. The concentrations are 1 M, 3 M and 6 M; 3 M is three times more concentrated than 1 M and 6 M is twice as concentrated as 3 M. Place a small piece of each metal which did **not** react with water in 1 mL of dilute (1 M) HCl in test tubes. Observe the rates of reaction (evolution of hydrogen gas in form of bubbles). Those metals which cause the evolution of hydrogen gas can be presumed to be more active than hydrogen.
Slow reactions often can be speeded by increasing the concentration of a reactant. If any of the metals reacts extremely slowly with the 1 M acid, or fails to react, test its reactivity in 3 M hydrochloric acid, HC1. Finally, try any metal in 6 M HC1 if it fails to react in 3 M HC1. Record in the table the data on the reactivity of each metal. Indicate if 3 M or 6 M HC1 was required and what you observed. Answer the questions following the table.

4. Reactions of Metals with Other Metal Ions

Compare the reactivity of iron, copper, and silver in the following ways, in separate test tubes place 5 mL of each of the specified salt solutions containing the metal cation, and add a piece of the indicated metal (a strip or a bright nail). Allow at least 15 minutes for a reaction to occur. The more active metal will replace the less active one from its solution. If a replacement reaction occurs you will easily recognize it by the deposition of the new metal on the nail or strip. The metals and salt solutions are:

a. iron (nail) in 0.1 M CuSO$_4$ solution
b. copper (strip) in 0.1 M FeSO$_4$ solution. This solution must be freshly made by the student. Weigh 0.13 grams of FeSO$_4$ and dissolve in 10 mL of water.
c. copper (strip) in 0.1 M AgNO$_3$ solution.
NAME: __________________________________________________________

DATA SHEET

1-3. Reactions of Metals with Oxygen, Water, and Acid

In the table below record your observations on the reactivity of the metals used in the experiment. Use qualitative terms such as fast(f), slow(s), vigorous(v), sluggish(sl), or no reaction(nr) to describe what happened. Make a separate record for each concentration of acid.

<table>
<thead>
<tr>
<th>Metal</th>
<th>With Air</th>
<th>Heated in Air</th>
<th>RXN WITH Water</th>
<th>RXN WITH 1M HCl</th>
<th>RXN WITH 3M HCl</th>
<th>RXN WITH 6M HCl</th>
</tr>
</thead>
<tbody>
<tr>
<td>Iron</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>tin</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>lead</td>
<td></td>
<td></td>
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<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>copper</td>
<td></td>
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<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>magnesium</td>
<td></td>
<td></td>
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<td></td>
<td></td>
</tr>
<tr>
<td>aluminum</td>
<td></td>
<td></td>
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<td></td>
<td></td>
</tr>
<tr>
<td>zinc</td>
<td></td>
<td></td>
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<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>sodium</td>
<td></td>
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</tr>
</tbody>
</table>

Did you observe any tarnishing on the surface of the freshly cut sodium? If so, what is the name of the corrosion product? Was sodium metal a reactant?

Write an equation for the reaction which occurred on the surface of the sodium metal.

Why is sodium normally stored under a liquid such as toluene or mineral oil?

Write equations for the reactions that occurred when you heated zinc and aluminum in the presence of oxygen.
What is the formula of the compound that was formed when magnesium was heated in air?

Make a sketch showing valence electrons for an atom of the free magnesium metal and for the magnesium ion in the resulting compound.

Did the atoms of magnesium lose electrons when the sample was heated?

Did other atoms accept electrons from magnesium? What kind of atoms?

Did metals react with oxygen more rapidly at flame temperature than at room temperature? Can you explain your observation?

Would any of the burning metals make a good night flare?

Would any of these metals be expected to react differently if the atmosphere were pure oxygen rather than the 20% oxygen in air? Explain.

Which type of iron sample has the greater surface area?

Which type of iron reacted more rapidly with oxygen? What is your evidence?

Write an equation showing the oxidation (rusting) of iron to ferric oxide.
Write equations for the reactions (if any) that occurred between hydrochloric acid and aluminum, and hydrochloric acid and copper.

Of the metals you have examined in this experiment, which would be the most suitable for use in jewelry? Which would be the least suitable?

Zinc is more active than copper. On the basis of this statement only, which of the following statements are true?

- Zinc has a greater tendency than copper to exist in the combined form (as a cation).
- It is easier to get copper metal from copper sulfate than to get zinc metal from zinc sulfate.
- Copper cyanide is more easily decomposed than zinc cyanide. Zinc tarnishes faster than copper.
- Zinc replaces hydrogen from acids but copper does not.

4. Reactions of Metals with Other Metal Ions

What, if anything, did you observe in each of the following trials? Write an equation for the reaction if one occurred.

Iron with CuSO₄ solution:

Copper with FeSO₄ solution:

Copper with AgNO₃ solution:

When a metal atom becomes an ion, or an ion becomes an atom, there is a transfer of electrons. In the preceding reactions, draw circles around the metal atoms which lost electrons and boxes around the metal ions which gained them (reactants).

Which is the more active: (a) iron or copper, (b) iron or silver, (c) silver or copper? Arrange the three metals in order of decreasing activity.

Arrange all of the metals tested in all parts of this experiment, including hydrogen, in order of decreasing activity.