in oxidation states. These concepts will be studied in greater detail in later experiments.

You should keep the three (related) definitions of equivalent mass in mind as you proceed:

- The mass of a metal that reacts with acid to produce ½ mole of $H_2(g)$
- The mass of a metal that reacts with acid to produce 11,200 mL of $H_2(g)$ at standard temperature (0 °C) and pressure (760 torr)
- The atomic mass of a metal divided by the change in charge that occurs in the reaction

Experimental Procedure

Special Supplies: Thermometer; pieces of metals (cut to size or issued as unknowns); 50-mL buret (for Method A only); fine copper wire; 25 × 250 mm test tube; two 500-mL Erlenmeyer flasks (for Method B).

Chemicals: Concentrated (12 M) hydrochloric acid, HCl; $Na_2CO_3(s)$.

SAFETY PRECAUTIONS: Concentrated hydrochloric acid is a lung irritant and causes skin and eye burns. Handle with care. Dispense it in a well-ventilated fume hood. Clean up any spills immediately. Protect your hands with plastic gloves.

NOTE The analysis samples for this experiment may be issued (1) as unknown metals, for you to calculate the equivalent mass; (2) as preweighed samples, for you to calculate and report the sample mass from the known equivalent mass; or (3) as Al–Zn alloys of different compositions, for you to calculate the percentage composition from the known equivalent masses. Prepare the units in your report accordingly, *before beginning the experiment.*

Your instructor will designate which of the preceding analyses and calculations you will do and which of the following alternate procedures you will follow. Method A is faster, but it limits you to smaller samples, to a fixed acid concentration, and to room temperature. Method B enables you to use larger samples and to control both temperature and acid concentration during the reaction.

NOTES TO INSTRUCTOR

Samples may be preweighed by a stock assistant on a rapid single-pan balance or, if they are in wire or ribbon form, cut to exact length to give a known mass and then be individually coded. The student may then report the corrected volume of hydrogen as a preliminary check on his or her work and finally report the calculated equivalent mass, or the mass of the sample from the known equivalent mass.

Obtain two samples of the metal to be used. Weigh these precisely on the analytical balance, at the same time taking care that the masses do not exceed the maximum permitted, so you do not generate more hydrogen than your apparatus can accommodate in either Method A or Method B. (For all laboratories except those at high elevation, maximum masses are as follows: for Method A, with a 50-mL buret, 0.12 g Zn, 0.032 g Al, 0.042 g Mg, 0.10 g Mn; for Method B, with a 500-mL flask, 1.10 g Zn, 0.40 g Mg, 0.30 g Al, 0.90 g Mn, 0.90 g Fe, 1.90 g Sn, 1.90 g Cd.)

Method A

Compress the weighed samples into compact bundles and wrap each sample in all directions with about 20 cm of fine copper wire, forming a small basket or cage, leaving 5 cm of the wire straight as a handle. This confines the particles as the metal dissolves and also speeds the reaction.[1]

Obtain and clean a 50-mL buret. Next measure the uncalibrated volume of the buret between the stopcock and the 50-mL graduation. Measure by filling the buret with water and draining the buret through the stopcock until the liquid level falls exactly to the 50-mL mark. Then use a 10-mL graduated cylinder to measure the volume delivered when the water level is lowered to the top of the stopcock. Using a funnel, pour into the buret the required amount of concentrated hydrochloric acid. Be careful not to allow the acid to touch your skin or clothing. Because of differences in the activity of the metals used, it is necessary to vary the amount of acid. For magnesium, use about 3 mL; for aluminum or zinc, use about 20 mL; for manganese, use about 7 mL.

[1]As the more active metal dissolves, it gives up electrons that move easily to the less reactive copper, where they react with hydrogen ions of the acid to form hydrogen gas. Note that the bubbles of gas form on the copper wire, thus keeping a larger surface of active metal exposed to the acid.

Fill the buret completely with water slowly and carefully, to avoid undue mixing of the acid. Insert the metal sample about 4 cm into the buret and clamp it there by the copper wire handle, using a one- or two-hole rubber stopper. Make certain no air is entrapped in the buret. Cover the stopper hole(s) with your finger[2] and invert the buret (Figure 13-1) in a 400-mL beaker partly filled with water. The acid, being more dense, quickly sinks, diffuses down the buret, and reacts with the metal. As the H_2 is generated, it collects at the top of the buret, expelling the HCl and water solution out the hole in the stopper at the bottom. (**Caution:** If the reaction is too rapid and the metal too close to the end of the buret, small bubbles of hydrogen may also escape from the buret as the acid solution is expelled. If so, repeat the experiment using less acid.)

After complete solution of the metal, let the apparatus cool to room temperature, since heat is generated by the reaction. Tap the apparatus to free any hydrogen bubbles adhering to the sides of the vessel or the copper wire. Measure the volume of gas liberated[3] and—without changing the position of the buret—measure the difference in height of the two water levels with a metric rule. Then calculate the equivalent pressure in millimeters of mercury (torr; mm Hg = mm solution × density of solution/density of mercury). Take the temperature of the gas by holding a thermometer in contact with the side of the buret. Raise the buret up out of the HCl solution in the beaker and allow the remainder of the solution to drain out of the buret into a large beaker. Slowly add solid sodium carbonate, Na_2CO_3, to the HCl solution until there is no further fizzing. Then flush the HCl solution down the drain with plenty of water and discard the copper wire in the wastebasket.

Obtain the barometer reading for the day. Repeat the determination with your second sample.

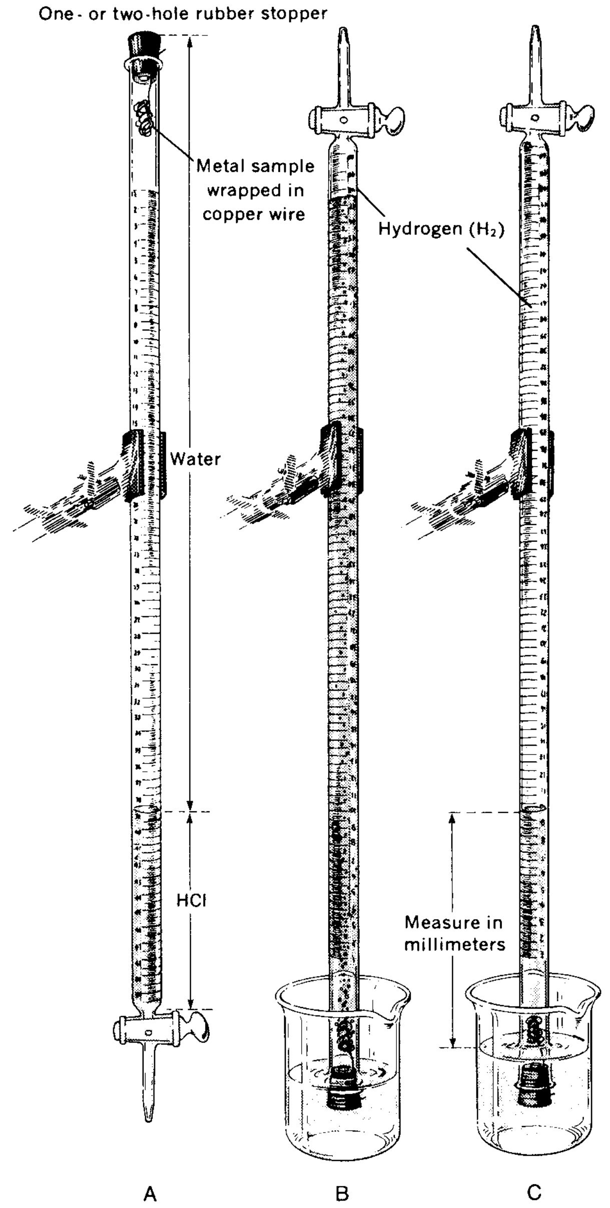

FIGURE 13-1 The appropriate volume of concentrated (12 M) HCl is added to the buret, and water is then layered on top of it until the buret is completely filled, as in A. Inversion of the buret in a beaker partly filled with water begins the reaction (B), which continues until all of the metal is gone and the buret is nearly full of hydrogen gas, as shown in C.

Method B

Set up the apparatus as sketched in Figure 13-2, utilizing a 25 × 250 mm test tube and a 500-mL flask to contain the evolved hydrogen. The exit tube C from the test tube, connected to the flask E, must not extend below either rubber stopper (so that gas will not be trapped). The longer glass tube in the test tube should extend nearly to the bottom and should be constricted to a small capillary and bent, as illustrated. Completely fill flask E with water and put a little water into flask F. Fill the siphon tube D, which extends to the bottoms of both flasks E and F, by blowing into the tube C.

Place the first carefully weighed sample of metal in the 25 × 250 mm test tube, as indicated. A fine copper wire, wrapped about the sample in all directions like a cage, as in Method A, may be of some

[2]Use a plastic glove for this step to ensure that no acid comes in contact with your skin.
[3]If a 50-mL buret is used, the volume of gas liberated equals 50 minus the final buret reading plus the volume of the uncalibrated portion of the buret.

EXPERIMENT 13
Report Guidelines

Calculations:

For each trial, calculate the following. Show complete calculations for one trial.

- Calculate pressure exerted by the water column in mmHg (*heights are inversely proportional to density*).

$$\frac{h_{Hg}}{h_{water}} = \frac{d_{water}}{d_{Hg}}$$

- Determine pressure of dry hydrogen gas collected (*gas pressure collected over water*)

$$P_{total} = P_{H2} + P_{water} + P_{column}$$

- Calculate moles of hydrogen gas produced (*ideal gas law, PV=nRT*).

- Calculate mass of metal reacted (*stoichiometry*)
  Write a balanced equation between the meal and HCl and determine the molar ratio of hydrogen gas to the metal

- Calculate the % of metal reacted.

$$\text{\% Metal reacted} = \frac{\text{mass of metal calculated}}{\text{mass of metal used in reaction}} \text{x100}$$

- Calculate volume of hydrogen gas produced at STP (*molar volume*)
  Using moles of hydrogen gas produced above calculate the volume at STP using molar volume.

- Calculate equivalent mass of metal (g metal/1 Eq):

$$\text{Equivalent mass of metal} = \frac{\text{mass of metal reacted}}{\text{Liter of } H_2 \text{ produced at STP}} \text{x11.2}$$

Questions

- Answer questions on the last page of report form.