

3. An inexpensive digital multimeter with a direct-current (DC) multirange ammeter provides a suitable readout. Two to four current ranges covering 50 microamperes (μA) to 200 or 300 milliamperes (mA) will work well.

4. To make a measurement, adjust the ammeter to the highest current range (200–300 mA). Put enough of the test substance in a 30-mL beaker to give a depth of about 1 cm. Lower the electrodes into the beaker so that they are immersed to a depth of 1 cm.

Then depress the push-button switch to get a current reading. Switch the ammeter to the most sensitive current range that can be used without driving the meter off scale. The solution conductivity is directly proportional to the current. Record the current reading, and return the ammeter to the highest current range.

5. Between each measurement, rinse the electrodes with distilled water and dry them with tissue. As an extra precaution, unplug the apparatus while cleaning the electrodes. Periodically clean the electrodes with steel wool and rinse with water.



WASTE COLLECTION:

Mercuric chloride, HgCl_2 , is toxic. After you have made measurements of the conductivity of 0.1 M HgCl_2 , pour the solution into a waste bottle. Do not pour it down the drain. Likewise, toluene and solutions of HCl (or acetic acid) in toluene should be poured into waste containers, not down the drain.

1. Electrolytes and Nonelectrolytes If you will be using the conductivity apparatus shown in Figure 18-1, read over the preceding notes describing how to use the apparatus. If you will use a commercial conductivity meter, follow your instructor's directions. When you are confident that you understand how to use the conductivity meter, measure and record the conductivity of the following *substances and/or their solutions in water*. Determine whether they are largely, moderately, or poorly ionized, or whether they are essentially nonelectrolytes. (For an apparatus like that shown in Figure 18-1, a nonelectrolyte will give a very small current, less than a few microamperes. A completely dissociated strong electrolyte, such as 0.1 M HCl, will typically give a current of more than 100 mA.) Test the pure substance first; then add some water to it and retest. In each

case, the magnitude of the current will be a measure of the conductivity of the solid or solution, indicating whether or not the substance contains ions that are free to move.

(a) To begin, test distilled (or deionized) water and tap water. What causes the difference in conductivity? Then test the following substances and their aqueous solutions: methanol (CH_3OH); glacial acetic acid (CH_3COOH); sucrose (table sugar, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$); sodium chloride (NaCl); and *dilute solutions only* (not the pure substances) of 0.1 M HCl, 0.1 M NaOH, and 0.1 M HgCl_2 . Be sure to dispose of the mercuric chloride, HgCl_2 , in the waste container reserved for this purpose. *Do not pour it down the drain—it is toxic.*

(b) Observe the *effect of the solvent* on the ionization of HCl in toluene and in water. First test a little pure toluene (C_7H_8) in a thoroughly dry beaker, and then a solution of HCl in toluene. (*Use the solution already on your reagent shelf, prepared by bubbling anhydrous HCl through dry toluene—do not add aqueous concentrated HCl to toluene.*) Then add 5 mL of deionized water to this solution, mix it well, and repeat the test with the electrodes immersed farther into the lower aqueous layer. (*If HCl in toluene is not available, you may substitute a 1 M acetic solution in toluene; the effect is less dramatic because acetic acid is not a strong acid in water like HCl.*) Be sure to dispose of the toluene solutions in the waste container reserved for this purpose.

(c) Compare the chemical behavior of 6 M HCl and 6 M acetic acid with your conductivity data. Put three or four marble chips into each of two clean beakers. Add 6 M HCl to one beaker and 6 M acetic acid to the other. Compare the rate of evolution of CO_2 gas. Put one or two pieces of mossy zinc metal into another pair of beakers, add 6 M HCl and 6 M acetic acid, and compare the rates of evolution of hydrogen (H_2) gas. Record your observations and write balanced chemical reactions for each of the four reactions you observed.

2. Typical Ionic Reactions You will experiment with some ionic reactions to determine the nature of their products. Through conductivity tests of the separate reactants and of the mixture after reaction, you will discover whether the acids, bases, and salts concerned are largely ionized (strong electrolytes) or only moderately ionized (weak electrolytes). For each reaction, write the total ionic equation and the

net ionic equation that best describes the reaction taking place. This will require careful thinking on your part. Use the following procedures.

(a) 0.01 M HCl with 0.01 M NaOH.⁴ These solutions can be prepared by diluting 5 mL of each of the 0.1 M solutions to 50 mL. It is an advantage to use both pairs of electrodes connected in parallel (see Figure 18-1) by placing one pair of electrodes in 0.01 M HCl and simultaneously placing the other pair in an equal volume of 0.01 M NaOH contained in a separate beaker. Then mix the solutions, divide them into two equal portions, and again place the two solutions simultaneously in contact with the two electrode pairs. (This compensates for the dilution effect of mixing two solutions.) Why is the conductivity of the product solution lower than that of the reactant solutions?

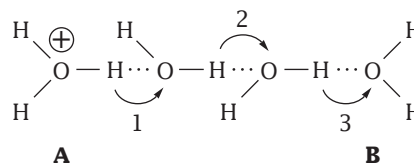
(b) 0.1 M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, with 0.1 M NH_3 .⁵ As in part 2(a), test equal volumes of the separate solutions, mix them, divide the solution into two equal parts, then remeasure the conductivity. Record your measurements.

(c) 0.1 M H_2SO_4 with 0.1 M $\text{Ba}(\text{OH})_2$. Measure and record the conductivity of each solution by itself. Then add two or three drops of 0.1% thymol blue indicator solution to the sulfuric acid solution, and add 0.1 M $\text{Ba}(\text{OH})_2$ solution one drop at a time to the sulfuric acid solution, stirring continuously. The sulfuric acid will be exactly neutralized when the indicator changes from its pink form through its yellow (acid) form to its blue (basic) form. If you add too much $\text{Ba}(\text{OH})_2$, overshooting the end point, add 0.1 M H_2SO_4 drop by drop, stirring continuously, until the *exact* end point—shown by a sharp color change of the indicator—is reached. Now, measure and record the conductivity of this mixture. Write balanced ionic equations to describe the reactions that take place.

CONSIDER THIS

Solutions of strong bases like NaOH and strong acids like HCl have much higher conductivity than a solution of the same concentration of NaCl. By examining a number of examples of this, scientists have discovered that it is the hydronium ion (H_3O^+) and hydroxide ion (OH^-) that are responsible for the higher conductivity.

Conductivity is related to the mobility of an ion (how fast it moves in solution under the influence of an electric field). In order to get from point A to point B in solution, most ions have to physically move from point A to point B, dragging along the cluster of water molecules that surrounds each ion. However, for hydronium ions, the whole hydronium ion does not need to move in order to transfer a positive charge from point A to point B. The way this works is shown in the following diagram:



Imagine that there is a hydronium ion at point A on the left end of a chain of water molecules. The water molecules are loosely associated by hydrogen bonding, as shown by the dotted lines representing the hydrogen bonds. The positive charge can be transferred from the hydronium ion at point A to the water molecule at point B by having three protons hop from one molecule to the next, as indicated by the three arrows. After transfer of the three protons, the water molecule at point B would be transformed into a hydronium ion (H_3O^+), so the charge associated with a hydronium ion would have moved from point A to point B without physical movement of the whole hydronium ion originally at point A. This rapid proton hopping (or proton transfer) process allows charge to move along much faster than physical movement of the whole ion, such as a Na^+ or Cl^- ion.

Now, you draw a similar diagram, starting with a hydroxide ion on the right at point B, that shows how the -1 charge associated with a hydroxide ion could move from point B to point A, by shifting protons left to right from one water molecule to the next. (Hydronium ions and hydroxide ions under the influence of the same electric field would move in opposite directions.)

⁴Quite dilute 0.01 M solutions are suggested so that you can interpret moderate conductivity changes more easily.

⁵Solutions in water are sometimes formulated as NH_4OH , called *ammonium hydroxide*. Because the evidence for the existence of such a molecule in solution is ambiguous, we will simply refer to a solution of ammonia in water as “aqueous ammonia solution.”