Conductivity Sensor Implementation

The circuit for the conductivity sensor is the same as for the photoresistor circuit studied earlier. A voltage divider circuit which includes the conductivity sensor and a 10kΩ resistor is energized by setting a digital output to HIGH. A portion of the 5V from the digital output is dropped across the conductivity sensor, and the remainder is dropped across the 10kΩ resistor:

$$\Delta V_{conductivity\ sensor} + \Delta V_{10k\Omega\ resistor} = 5\ volts$$

Figure 1 shows a diagram of the circuit. An analog input on the Arduino is used to measure the voltage drop across the 10kΩ resistor. While we could just as easily measure the voltage drop across the conductivity sensor by making it the "resistor" closest to ground, configuring the circuit as shown causes the voltage measured at the analog input to <u>increase</u> as the water becomes more salty. Configuring the circuit so that more salt causes a higher system output makes analysis of system output more intuitive.
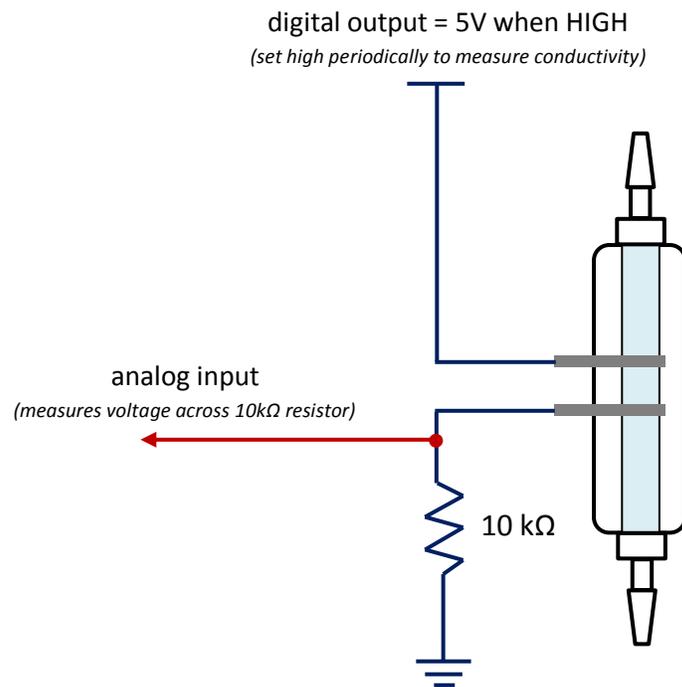
Figure 1 – Diagram of the conductivity sensor circuit.

From the diagram above, it is clear that charge must somehow enter and leave the salt water if current is to flow through the circuit. The answer to how this occurs lies in electrochemistry. Salts, like NaCl, consist of tightly bonded ions. When salts are placed in water, the ionic bonds are weakened, and the ions become mobile. The ions are surrounded by the polar water molecules (they are hydrated). Figure 2 depicts Na^+ and Cl^- ions surrounding the electrodes of the conductivity sensor when no voltage is applied across the electrodes.
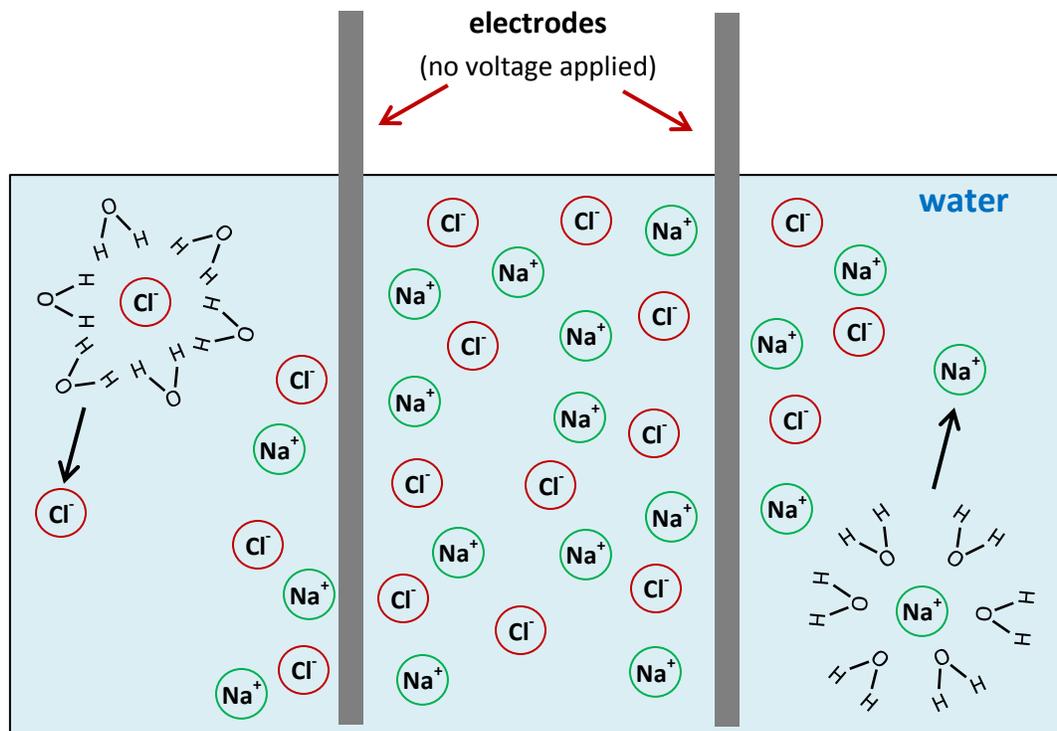
Figure 2 – Hydration of NaCl in water.

When current passes through the conductivity circuit, electrons do not flow directly flow through the water between the electrodes in the same way as electrons flow through a wire. Instead, when a sufficient voltage is placed across the electrodes, one of the electrodes will become positively charged and the other will become negatively charged, as depicted in Figure 3.

The positively charged electrode is called the **anode**; the anode extracts electrons from the solution as part of an **oxidation** reaction:

$$2\,Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$$
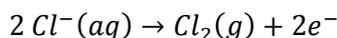
Notice that the Cl⁻ ions lose electrons; these electrons will be conducted away through the electrode and on through the connecting wire as they migrate toward the 5V source (provided by the digital output).

The negatively charged electrode becomes the cathode; the cathode draws electrons from the ground and provides them for a **reduction** reaction:

$$2\,H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$$
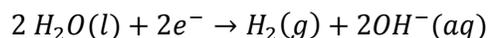
While it is intuitive to expect the Na$^+$ ion to be reduced at the cathode, this is not the dominant reaction at the cathode as we will see later. Instead, water is reduced while Na$^+$ serves as a spectator ion (although some Na$^+$ ions may combine with the OH⁻ ions to produce NaOH). That is, it is more

energetically favorable for the water to be reduced to form H_2 and OH^- than it is for Na^+ to accept the electron to become an uncharged Na atom.
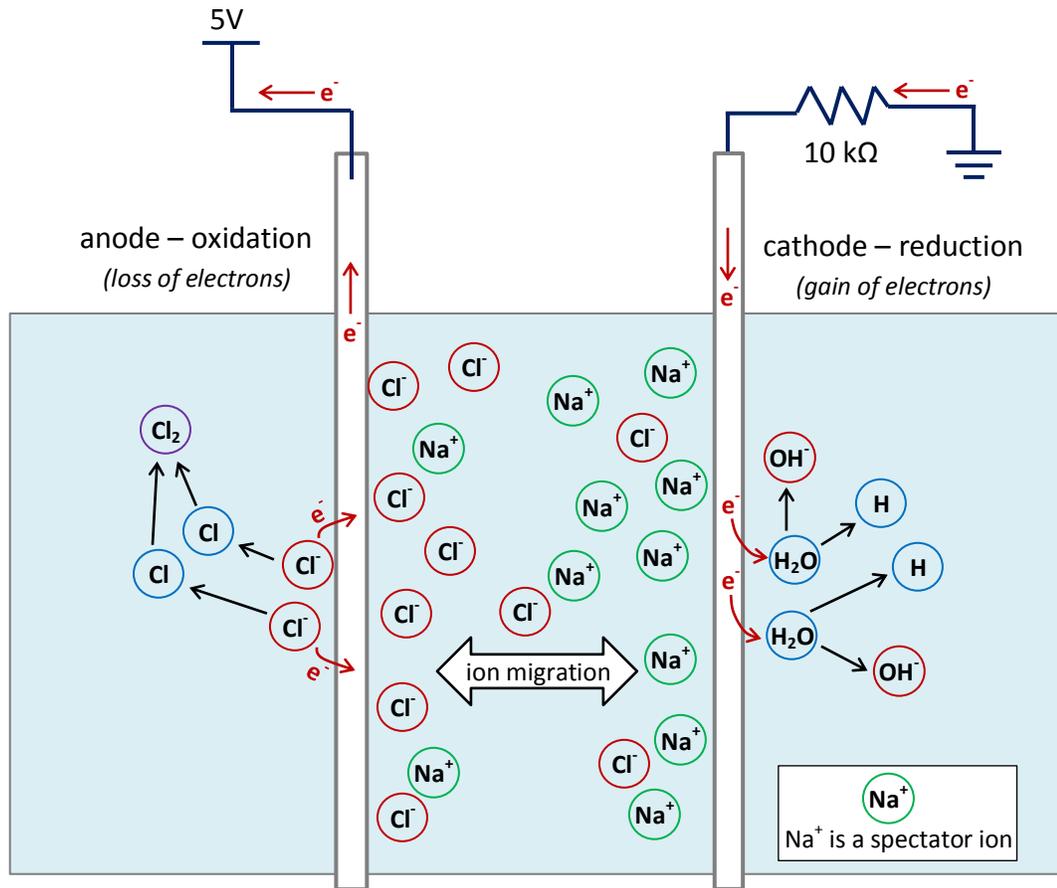
Figure 3 – Depiction of reactions that occur when conductivity sensor is energized.

Why is H_2O reduced and not Na^+ ions? To understand why H_2O is reduced in favor of Na^+ ions, we use the "table of standard reduction potentials," as shown in Table 1. Reactions that are higher in the table will induce a higher voltage when connected to a reference half-cell, as shown in Figure 4. A reference half-cell is a platinum electrode immersed in a 1.0M solution of H^+ ions. Half-reactions higher up in Table 1 will create a positive voltage when connected to the standard half-cell, and half-reactions lower in the table will create a negative voltage when coupled with the standard half-cell.

Table 1 – Standard reduction potentials for a few aqueous solutions (more complete tables exist).

	Electrode Reduction Half-Reaction	**Voltage Output** *(when electrode is coupled with a 1.0M H+ ion solution using a platinum electron)*
increasingly inert	$F_2(g) + 2e^- \rightarrow 2F^-(aq)$	+2.87 V
	$H_2O_2(aq) + 2H^+ + 2e^- \rightarrow 2H_2O(l)$	+1.78 V
	$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$	+1.52 V
	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36 V
	$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$	+0.40 V
	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.34 V
	$2H^+ + 2e^- \rightarrow H_2(g)$	0.00 V
	$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$	-0.25 V
	$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	-0.83 V
	$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	-1.66 V
increasingly active	$Na^+(aq) + e^- \rightarrow Na(s)$	-2.71 V
	$Li^{-1}(aq) + e^- \rightarrow Li(s)$	-3.04 V

Figure 4 – Standard hydrogen half-cell coupled with another half-cell (see Table 1 for other half-cells).

It is important to understand that coupling some half-cells together will create a positive voltage and thereby produce electric current (like a battery). Other half-cell combinations will combine to form a negative voltage, and for these reactions to occur, an external voltage must be supplied across the electrodes.

We can think of the conductivity sensor as two half-cell reactions coupled together. The three primary options for electrode reactions are shown in Table 2.

Table 2 – Primary reactions that can occur at electrodes of conductivity sensor.

	Possible Half-Reaction	Reference Voltage
1	$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	-0.83 V
2	$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$	-1.36 V *(direction & sign swapped)*
3	$Na^+(aq) + e^- \rightarrow Na(s)$	-2.71 V

If half-cells for reactions 1 and 2 in the table above are coupled, the net reaction and the resulting voltage are shown in Table 3.

Table 3 – Net reaction for oxidizing Cl⁻ and reducing water.

anode	$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$	-1.36 V
cathode	$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	-0.83 V
net reaction	$2Cl^-(aq) + 2H_2O(l) \rightarrow Cl_2(g) + H_2(g) + 2OH^-(aq)$	**-2.19 V**

Coupling half-cells for reactions 1 and 3 in Table 2 together will result in a voltage of -3.54V; likewise, coupling half-cells for reactions 2 and 3 together will result in a voltage of -4.07V. Since reactions 1 and 2 require the least amount of external voltage to initiate the reaction (-2.19V), this is the reaction that is most likely to occur.

Although our half-cells are not "standard half-cells" since the concentrations of ions are not set to 1.0M, the idea that we are able to determine the most probable oxidation and reduction reactions based on standard reduction potential is useful for understanding what physically occurs when we energize our conductivity sensor. Table 3 shows that the net reaction will require a voltage greater than 2.19V to drive the reaction and to get any electrical current at all; our conductivity sensor may not work if we apply less than this voltage. The voltage required to drive the circuit will likely depend strongly on the concentration. The voltages for the other potential oxidation / reduction reaction combinations in Table 2 (-3.54V and -4.07V) reveals that we can have multiple reactions occurring simultaneously if we apply enough voltage.