Instrumentation
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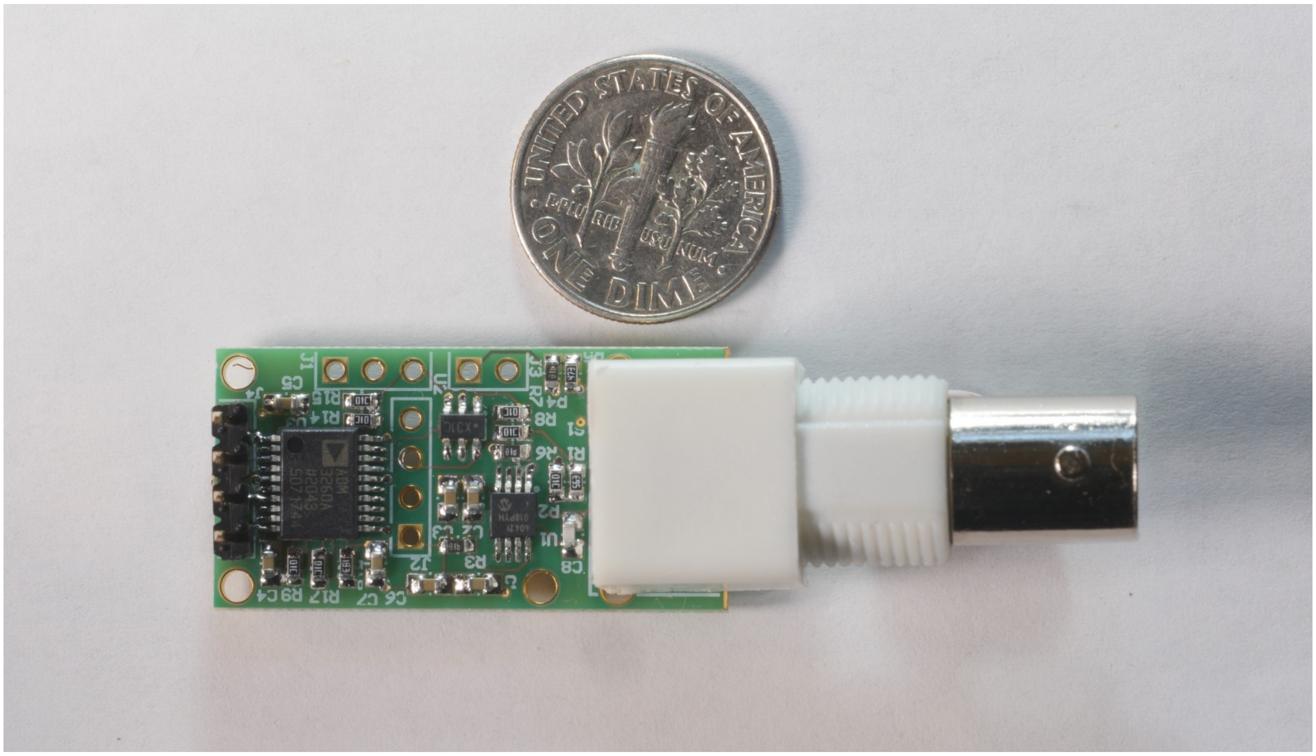
InstrumentationHub cAFE1 Analog Front End Board



cAFE1 is a specially designed front end to handle electrochemical measurement cells. A high impedance instrumentation amplifier with adjustable gain and offset is connected to a 12bit analog to digital converter [1]. The board electronics is fully electrically isolated to prevent cross talk and leakage currents from interfering with measurements. Controller connection is via an i2c interface [2].

An **electrochemical measurement cell** is a device that generates an electrical potential depending on the ion concentration of solutes. There are different types to measure different types of ions.

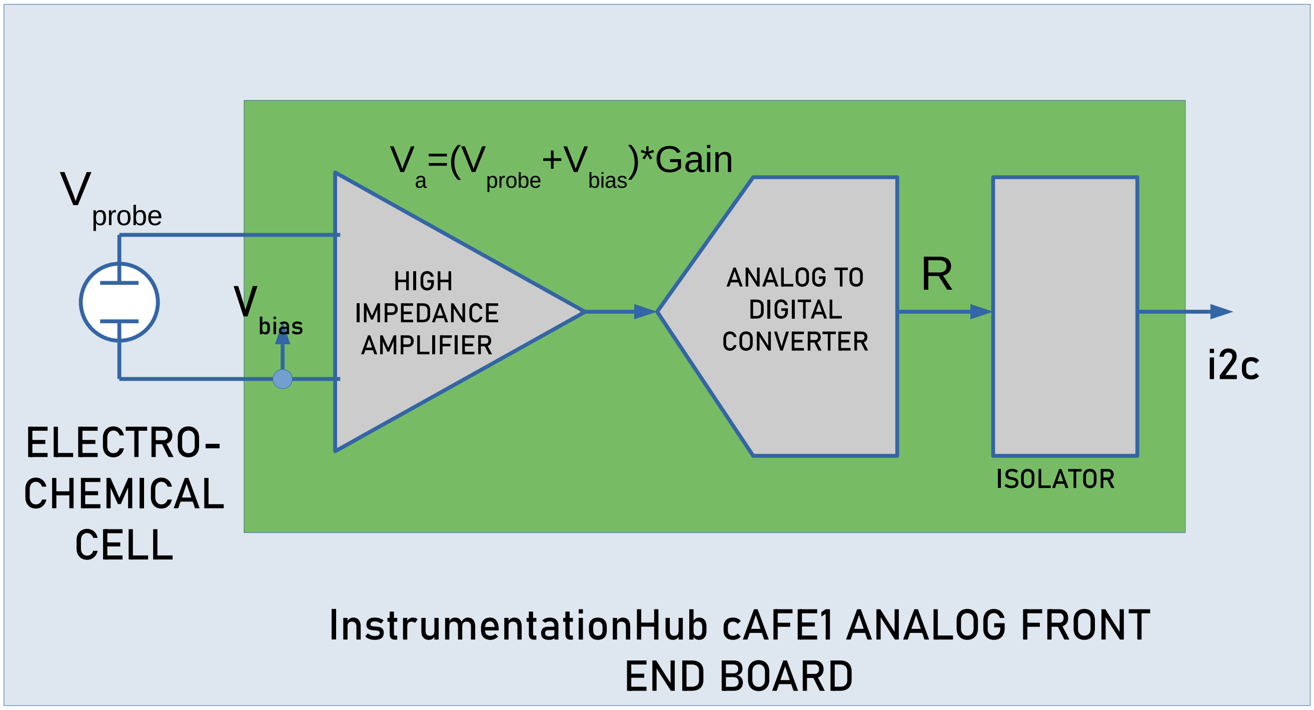
Some electrochemical cells include sensors for:
- pH
- ORP (oxygen reduction potential)
- Dissolved Oxygen
- Ammonia
- Nitrite and Nitrates

Some main applications for these sensors include:

- nutrient management for hyroponic/aeroponic agricultural systems
- pool water management
- aquarium, aquaculture management

Different electrochemical cells can have different electrical requirements that may need minor adjustments to the cAFE1 board.

cAFE1 supports all pH cells in its default configuration



cAFE1 Block Diagram

pH Calculations

A pH electrode measures hydrogen ion (H+) activity and produces an electrical potential or voltage. The operation of the pH electrode is based on the principle that an electric potential develops when two liquids of different pH come into contact at opposite sides of a thin glass membrane. pH measurement is the measurement of this potential difference and conversion to pH units.

Note: Units are designated in bold font. i.e. **°C, °F, V, pH, count**

Theory

pH in an aqueous solutions is a measure of hydrogen and hydroxide ions.

Water molecules dissociate in hydrogen (H+) and hydroxide (OH–) ions.

The concentration (mol/L) of hydrogen ions [H+] multiplied by the concentration (mol/L) of hydroxide ions
[OH–] is constant and is called Kw, the dissociation constant for water.

Kw depends on temperature and is 1.008×10^{-14} at 25**°C**. This means that the concentrations of [H+] and [OH–] are each 1×10^{-7} **mol/L** at 25**°C**. This is why pH is 7 in neutral water where [H+] and [OH-] are equal as we will see when we see the definition of pH.

Temperature °C	**Kw**
10	0.2920×10^{-14}
15	0.4505×10^{-14}
20	0.6809×10^{-14}
25	1.008×10^{-14}
30	1.469×10^{-14}

Acids in water increase the [H+] and because the product [H+] [OH–] must be constant, decrease the [OH–]. Bases increase [OH–] and decrease [H+].

pH is defined, by the Sörenson Equation, as the negative logarithm of the H+ concentration in a given solution:

$$pH = -log[H+]$$

Example:

$$[H+] = 1 \textbf{ mol/L} = 10^{0} \rightarrow pH = 0 \text{ (ACIDIC)}$$

$$[H+] = 10^{-7} \textbf{ mol/L} = 10^{7} \rightarrow pH = 7 \text{ (NEUTRAL)}$$

$$[H+] = 10^{-14} \textbf{ mol/L} \rightarrow pH = 14 \text{ (ALKALINE)}$$

If the H+ ion concentration changes by the factor of ten, the pH value changes by one unit. This illustrates how important it is to be able to measure the pH value to a tenth of a unit (or even a hundredth of a unit in critical cases).

The Nernst Equation

The general mathematical description of electrode behavior was described by the 19th century German chemist, Hermann Walther Nernst (1864 – 1941). He introduced the Nernst equation in 1889, expressed as:

$$E_{nernst} = 2.3 * (RT/nF)$$

where:

E_{nernst} = total potential difference
R = universal gas constant (in ***Joules/mol-Kelvin****) (8.31439 J ** $\mathbf{mol^{-1}}$ **x** $\mathbf{K^{-1}}$*)*
T = absolute temperature (in ***Kelvin****)*
n = charge of the ion (in this case $n_H = 1$)
F = Faraday constant (in ***Coulombs/mol****) (96495.7 C ** $\mathbf{mol^{-1}}$*)*

The entire term "2.3RT/nF" is called the Nernst factor, or slope factor. This term provides the amount of change in total potential for every ten-fold change in ion concentration. For hydrogen ion activity, where n=1, the Nernst factor is 59.16 **mV** for every ten-fold change in activity at 25**°C**. This means that for every pH unit change, the total potential will change 59.16 **mV**. This derived as follows:

$$E_{nernst} = 2.3 * ((8.31439*1000\,\mathbf{mV/V})/96495.7)Tk$$
$$= 0.19818 * Tk = 0.19818 * (273.15 + Tc)$$

For 25 **°C**

$$E_{nernst} = 0.19818 * (273.15 + 25) = 59.1\ \mathbf{mV/pH}$$

For an ideal probe, voltage measured across electrodes V_{probe} for different pH values at 25 **°C** will be as follows

pH	V_{probe} **(mV)**
.	..
5	118.32
6	59.16
7	0
8	-59.16
9	-118.32
.	..

Measuring the pH electrode potential

The pH measuring electrode and the pH reference electrode form a so-called pH measuring chain within the measured medium. This chain can be compared to a battery of which the voltage produced depends on the measured medium.

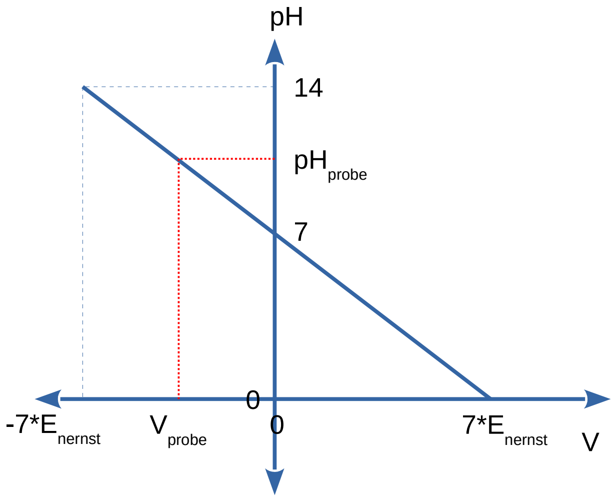
The difference in potential between the measuring electrode and the reference electrode is a function of the pH value of the measured medium. In theory the voltage changes by 59.16 **mV** per pH unit at 25 **°C** according to the Nernst equation. The voltage produced by the pH measuring chain is large enough not to present any problem for a measurement. But the measuring chain is a voltage source from which no current can be drawn, not even the low current which a moving coil of a DC voltmeter draws. The potential difference of the measuring chain has to be measured without drawing any current from the voltage source, otherwise the voltage would be reduced and the pH measurement would be drastically falsified. The reason is the high electrical resistance of the glass electrode which is mainly determined by the resistance of the glass membrane.
The resistance values of a glass membrane vary between 10 **MΩ** and 1 000 **MΩ** at 25 **°C** and increase 10 times at a temperature decrease of 25 **°C**.

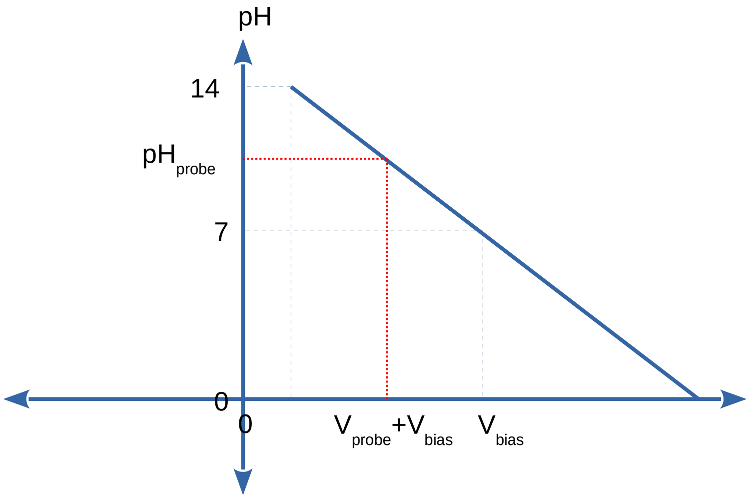
The upper limit of the membrane resistance lies at 5 **GΩ** (5 x 10^9 **Ω**). Membrane resistance that is too high causes faults and disturbances in the electronic measuring instrument. The e.m.f. produced by the high resistance measuring chain can only be measured by an instrument having such a high internal resistance that it does not draw a current from the chain. For practical purposes the pH meter or pH transmitter should therefore have an internal resistance of at least 10^{12} Ohms (1**TΩ**).

Using cAFE1 for pH Measurements

There are 3 transformations:

1) A bias voltage is applied to move the voltage range from -/+ to fully + because the amplifier has a single positive power supply. The probe voltage is now $V_{probe} + V_{bias}$.





2) The signal is amplified by the Gain factor of the amplifier. The output of the amplifier is given by.

$$Va = Gain *(V_{probe} + V_{bias})\ \mathbf{V}$$

The pH/Voltage relationship is linear equation of the form:

$$pH = a + b_{probe}V_a$$

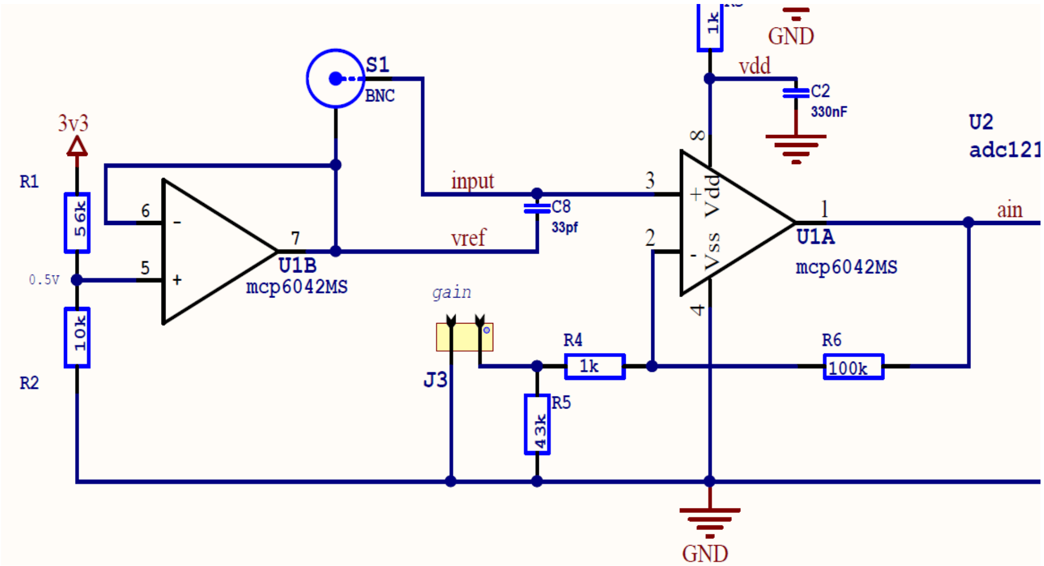
3) The signal Va is digitized by the ADC resulting in a digital **count** value R

$$R = \frac{2^{12}\,\boldsymbol{count}}{3.3\,\boldsymbol{V}}\cdot V_a \quad = \left(\frac{2^{12}\,\boldsymbol{count}}{3.3\,\boldsymbol{V}}\cdot(V_{probe}+V_{bias})\cdot Gain \right) \textbf{count}$$

The pH/count relationship is linear equation of the form:

$$pH = A + BR \qquad (0)$$

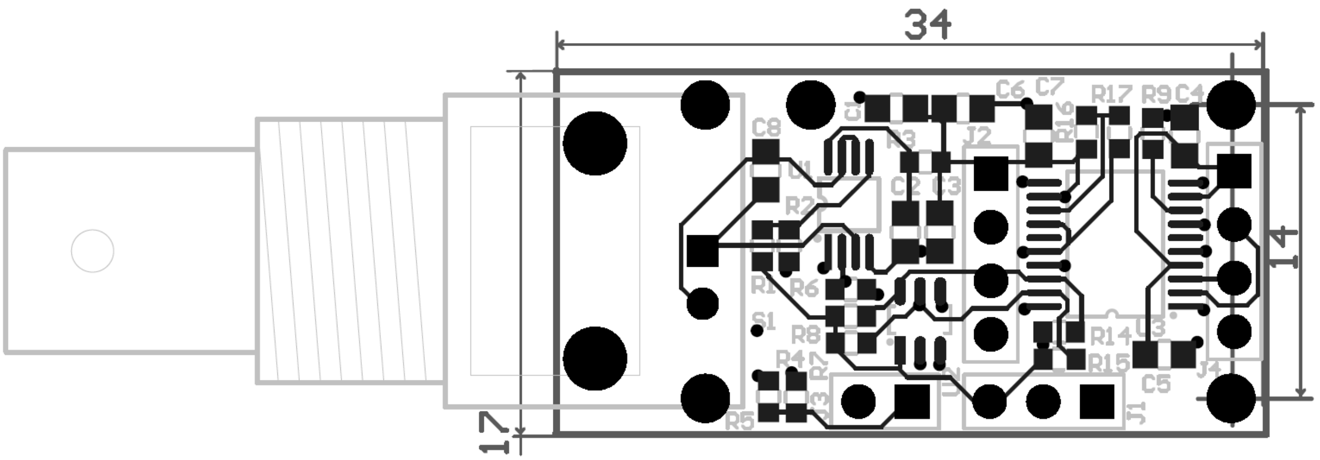
V_{bias} and Gain



$$V_{bias} = \frac{R_2}{(R_1 + R_2)} \cdot 3.3\,\boldsymbol{V} = \frac{10}{66} \cdot 3.3\,\boldsymbol{V} = 0.5\,\boldsymbol{V}$$

$$Gain = 1 + \frac{R_6}{(R_4 + R_5)} = 1 + \frac{100}{44} = 3.27$$

Physical Dimensions



[1] https://www.ti.com/product/ADC121C027

[2] cAFE1 i2c interface