A study on connection between chemistry and electricity

K.A. Khan[1] and Md. Abdul Awal[2]
[1]Department of Physics, Jagannath University, Dhaka-1100, Bangladesh
[2]Department of Chemistry, Hasadah model kamil madrasha, Jibannagar, Chuadanga

ABSTRACT

There is a good connection between chemistry and electricity from an ancient time. One of the great scientists named Alessandro Volta has discovered Voltaic cell in 1973. It was a great discovery for the human being. Nicholson and Carlise showed that electric current could decompose water into hydrogen and oxygen in 1800. This experiment was the most important significant experiment. This implied that the atoms of oxygen and hydrogen were associated with negative and positive electric charges. Finally it was proved that these positive and negative charges were the source of the bonding forces between them. One of the Swedish chemists named Berzelius proposed that all atoms are electrified and among them hydrogen & metals are positive and the nonmetals are negative in 1812. The applied voltage for electrolysis was thought to overpower the attraction between these opposite charges, pulling the electrified atoms apart in the form of ions. After about one hundred years later he gave a electron pair theory. Lewis could offer a significant improvement over this view of chemical bonding. Meanwhile the use of electricity as a means of bringing about chemical change continued to play a central role in the development of chemistry. Humphrey Davey prepared the first elemental sodium by electrolysis of a sodium hydroxide melt. It was left to Davey's former assistant, Michael Faraday, to show that there is a direct relation between the amount of electric charge passed through the solution and the quantity of electrolysis products. James Clerk Maxwell immediately saw this as evidence for the "molecule of electricity", but the world would not be receptive to the concept of the electron until the end of the century.

Keywords: *Electro neutrality, Ions, Voltaic cell, Galvanic cell.*

I. Introduction:

Nature seems to strongly discourage any process that would lead to an excess of positive or negative charge in matter. For example, we immerse a piece of zinc metal in pure water. A small number of zinc atoms go into solution as Zn ions, leaving their electrons behind in the metal: $Zn(s) = Zn^{2+} + 2e^-$. As this process goes on, the electrons which remain in the zinc cause a negative charge to build up within the metal which makes it increasingly difficult for additional positive ions to leave the metallic phase. A similar buildup of positive charge in the liquid phase adds to this inhibition. Very soon, therefore, the process comes to a halt, resulting in a solution in which the concentration of Zn^{2+} is still too low (around 10–10 M) to be detected by ordinary chemical means.

There would be no build-up of opposing charges in the two phases if the excess electrons could be removed from the metal or the positive ions consumed as the reaction proceeds. For example, we could drain off the electrons left behind in the zinc through an external circuit that forms part of a complete electrochemical cell; this we will describe later. Another way to remove electrons is to bring a good electron acceptor (that is, an oxidizing agent) into contact with the electrode. A suitable electron acceptor would be hydrogen ions; this is why acids attack many metals. For the very active metals such as sodium, water itself is a sufficiently good electron acceptor. The degree of charge unbalance that is allowed produces differences in electric potential of no more than a few volts, and corresponds to unbalances in the concentrations of oppositely charged particles that are not chemically significant. There is nothing mysterious about this prohibition, known as the electro neutrality principle; it is a simple consequence of the thermodynamic work required to separate opposite charges, or to bring like charges into closer contact. The additional work raises the free energy of the process, making it less spontaneous. The only way we can get the oxidation of the metal to continue is to couple it with some other process that restores electro neutrality to the two phases. A simple way to accomplish this would be to immerse the zinc in a solution of copper sulfate instead of pure water. As you will recall if you have seen this commonly-performed experiment carried out, the zinc metal

quickly becomes covered with a black coating of finely-divided metallic copper. The reaction is a simple oxidation-reduction process, a transfer of two electrons from the zinc to the copper: $Zn(s) = Zn^{2+} + 2e^{-}$, $Cu^{2+} + 2e– = Cu(s)$
The dissolution of the zinc is no longer inhibited by a buildup of negative charge in the metal, because the excess electrons are removed from the zinc by copper ions that come into contact with it. At the same time, the solution remains electrically neutral, since for each Zn ion introduced to the solution, one Cu ion is removed. The net reaction: $Zn(s) + Cu^{2+} = Zn^{2+} + Cu(s)$ quickly goes to completion.

II. Voltage differences at interfaces

The transition region between two phases consists of a region of charge unbalance known as the electric double layer. As its name implies, this consists of an inner monomolecular layer of adsorbed water molecules and ions, and an outer diffuse region that compensates for any local charge unbalance that gradually merges into the completely random arrangement of the bulk solution. In the case of a metal immersed in pure water, the electron fluid within the metal causes the polar water molecules to adsorb to the surface and orient themselves so as to create two thin planes of positive and negative charge. If the water contains dissolved ions, some of the larger (and more polarizable) anions will loosely bond (chemisorb) to the metal, creating a negative inner layer which is compensated by an excess of cations in the outer layer. Electrochemistry is the study of reactions in which charged particles (ions or electrons) cross the interface between two phases of matter, typically a metallic phase (the electrode) and a conductive solution, or electrolyte. A process of this kind can always be represented as a chemical reaction and is known generally as an electrode process. Electrode processes take place within the double layer and produce a slight unbalance in the electric charges of the electrode and the solution. Much of the importance of electrochemistry lies in the ways that these potential differences can be related to the thermodynamics and kinetics of electrode reactions. In particular, manipulation of the interfacial potential difference affords an important way of exerting external control on an electrode reaction. The interfacial potential differences which develop in electrode-solution systems are limited to only a few volts at most. This may not seem like very much until you consider that this potential difference spans a very small distance. In the case of an electrode immersed in a solution, this distance corresponds to the thin layer of water molecules and ions that attach themselves to the electrode surface - normally only a few atomic diameters. Thus a very small voltage can produce a very large potential gradient. For example, a potential difference of one volt across a typical 10–8 cm interfacial boundary amounts to a potential gradient of 100 million volts per centimeter - a very significant value indeed! Interfacial potentials are not confined to metallic electrodes immersed in solutions; they can in fact exist between any two phases in contact, even in the absence of chemical reactions. In many forms of matter, they are the result of adsorption or ordered alignment of molecules caused by non-uniform forces in the interfacial region. Thus colloidal particles in aqueous suspensions selectively adsorb a given kind of ion, positive for some colloids, and negative for others. The resulting net electric charge prevents the particles from coming together and coalescing, which they would otherwise tend to do under the influence of ordinary van der Waals attractions.

III. Interfacial Voltage differences are not directly observable.

The usual way of measuring a voltage difference between two points is to bring the two leads of a voltmeter into contact with them. It's simple enough to touch one lead of the meter to a metallic electrode, but there is no way you can connect the other lead to the solution side of the interfacial region without introducing a second electrode with its own interfacial potential, so you would be measuring the sum of two voltage differences. Thus single electrode voltage, as they are commonly known, is not directly observable.

IV. Electrochemical cells

Although it is physically impossible to measure or manipulate the potential difference between a piece of metal and the solution in which it is immersed, we can easily measure a potential difference between two such electrodes immersed in a solution. The result will be the sum of the two electrode potentials, we shall see farther on that such measurements can be supply all the information we need in order to characterize the two electrode reactions. The arrangement is called a galvanic cell. A typical cell might consist of two pieces of metal, one zinc and the other copper; each immersed each in a solution containing a dissolved salt of the corresponding metal. The two solutions are separated by a porous barrier that prevents them from rapidly mixing but allows ions to diffuse through. If we simply left it at that, no significant amount of reaction would take place. However, if we connect the zinc and copper by means of a metallic conductor, the excess electrons that remain when Zn^{2+} ions go into solution in the left cell would be able to flow through the external circuit and into the right electrode, where they could be delivered to the $Cu2^{+}$ ions which become "discharged", that is, converted into Cu atoms at the surface of the copper electrode. The net reaction is the same as before - the oxidation of zinc by copper(II) ions: $Zn(s) + Cu^{2+} = Zn^{2+} + Cu(s)$ but this

time, the oxidation and reduction steps take place in separate locations: left electrode: $Zn(s) = Zn^{2+} + 2e^-$ oxidation right electrode: $Cu^{2+} + 2e^- = Cu(s)$ reduction

IV. Electrochemical cells allow measurement and control of a redox reaction.

The reaction can be started and stopped by connecting or disconnecting the two electrodes. If we place a variable resistance in the circuit, we can even control the rate of the net cell reaction by simply turning a knob. By connecting a battery or other source of current to the two electrodes, we can force the reaction to proceed in its non-spontaneous or reverse direction. By placing an ammeter in the external circuit, we can measure the amount of electric charge that passes through the electrodes, and thus the number of moles of reactants that get transformed into products in the cell reaction. Electric charge q is measured in coulombs. The amount of charge carried by one mole of electrons is known as the faraday, which we denote by F. Careful experiments have determined that 1 F = 96467 c. For most purposes, you can simply use 96,500 coulombs as the value of the faraday

V. Results and Discussions

Consider in an electrochemical cell of how much mass would the zinc electrode lose if a current of 0.15 amps flows through the external circuit for 1.5 hours? The amount of charge passing between the electrodes is (0.15 amp) ´ (54600 sec) = 810 c or (810 c) / (96500 c F^{-1}) = 0.0084 F. Since the oxidation of one mole of Zn to Zn^{2+} results in the removal of two moles of electrons, the number of moles of Zn removed from the electrode is 0.0042, corresponding to a weight loss of (.0043 M) ´ (65.37 g M^{-1}) = 0.275 g .

VI. Conclusion

There are several other conventions relating to cell notation and nomenclature that you are expected to know:

• The anode is where oxidation occurs, and the cathode is the site of reduction. In an actual cell, either electrode can have either identity, depending on the direction in which the net cell reaction is occurring.

• If electrons flow from the left electrode to the right electrode (from cathode to anode) when the cell operates in its spontaneous direction, the potential of the right electrode will be higher than that of the left, and the cell potential will be positive.

• "Conventional current flow" is from positive to negative, which is opposite to the direction of the electron flow. This means that if the electrons are flowing from the left electrode to the right, a galvanometer placed in the external circuit would indicate a current flow from right to left.

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