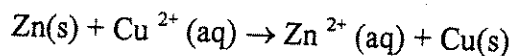


ELECTROCHEMICAL CELLS

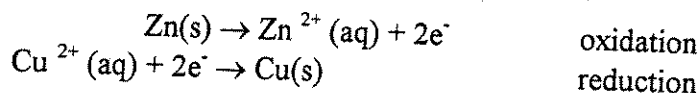
Theory

Electrochemistry is the detailed study of electron-transfer (or oxidation/reduction, redox) reactions. It is a very wide field of endeavor, covering such subjects as batteries, corrosion and reactivities of metals, and electroplating. Oxidation-reduction reactions find their most important use in the construction of voltaic (galvanic) cells (chemical batteries).

Consider the following reaction:



In this process, metallic elemental zinc has been added to a solution of dissolved copper(II) ions. A reaction occurs. The metallic zinc dissolves, producing a solution of zinc ion. At the same time, metallic elemental copper forms from the copper(II) ions that are present in solution. To see what really is happening in this reaction, it is helpful to split the given equation into two half-reactions:



The zinc half-reaction is called an oxidation half-reaction. Oxidation is a process in which a species **loses electrons** to some other species. In this half-reaction, metallic zinc loses two electrons in becoming zinc(II) ions. The copper half-reaction is called a reduction half-reaction. Reduction is a process in which a species **gains electrons** from some other species. In the copper half-reaction, copper(II) ions gain two electrons (i.e. the electrons that had been lost by the zinc atoms) and become metallic copper.

An oxidation-reduction reaction can be considered to be the sum of two half-reactions, one involving the loss of electrons from an element (oxidation) and the other involving the gain of electrons by an element (reduction). We can assign to each half-reaction a potential that indicates the relative driving force for that half-reaction. For many oxidation-reduction reactions, it is possible to set up an electrical cell in which the two half-reactions occur at separate electrodes and from which the overall cell potential is the sum of the potentials of the half-reactions. In the case of the copper and the zinc half-cells we can set up the reaction so that the zinc metal and copper(II) ion solution are physically separated (in separate beakers, for example) but are connected electrically by a conducting wire. Since the reaction that occurs is a transfer of electrons, this can now occur through the wire, thereby producing an electrical current. We could place a motor or light bulb along the wire joining the copper/zinc beakers and make use of the electrical current produced by the reaction. We have constructed a **battery** (or voltaic cell) consisting of a zinc half-cell and a copper half-cell. In this battery, the electrode at which oxidation occurs is called the anode and the electrode at which the reduction reaction takes place is called the cathode. A second connection will have to be made between the two beakers to complete the electrical circuit. This is done with a salt bridge which allows ions to flow in order to keep the electrical neutrality of the system.

Electrolysis is the use of an electrical current to force a chemical reaction to occur that would ordinarily not proceed spontaneously. In an electrolytic cell an electrical current from an external source is passed through a solution, causing a redox reaction to occur. At the positive electrode (also called the anode), an oxidation half reaction takes place. At the negative electrode (cathode), a reduction half reaction takes place.

According to Faraday's laws, the amount of substance produced or consumed in electrolysis is proportional to the amount of electricity that passes through the cell. The electrical charge on one mole of electrons is called a Faraday. The quantity of charge is generally measured in Coulombs. One Faraday is equal to 96,485 Coulombs.

The electrical current in amperes (amps) gives the number of coulombs of charge per second passing through the cell. The total charge used is given by the equation:

$$q = i \times t$$

where q is the total charge in coulombs, i is the current passed through the system in amps and t is the time in seconds that this current was passed through the cell.

Experimental Procedure

Faraday and Avogadro: Electrolysis

Obtain two pieces of copper wire and clean them with steel wool to remove the processing wax. Make sure that you keep fingerprints to a minimum. Wax and skin oil will act as an insulating barrier. Rinse them with acetone and weigh them. Bend each wire of your pair in a way you can identify (one with a hook and the other with a right angle bend).

Add around 80 mL of a 0.5 M CuSO_4 solution into a 150 mL beaker. Put your two wires inside the beaker (on opposite sides) and use a rubber stopper to keep them apart and in place. Connect your system (the beaker with the Cu wire electrodes and the copper sulfate solution) to a power supply as instructed by your professor. (3 or 4 groups will share an apparatus as shown in figure 1 and must cooperate in setting it up. Use jumper wires to connect the copper electrodes between the groups.)

When the groups are all ready, read the time and then have one person turn the power supply on and adjust the current to ~ 0.2 A to start the electrolysis. STIR! Read the ammeter as accurately as you can every minute while stirring the solution; the current will not necessarily be constant. Let the experiment run for 20 minutes and record the ending time. Carefully remove the electrodes, wash them gently with water and acetone, and let them dry a few minutes. Inspect the CuSO_4 solution. Did it change color (or intensity)? Why or why not? Is there any suspended matter in the beaker? If so, what is it and where did it come from? Was it present at the beginning?

Weigh each electrode accurately. Does the loss of weight of one electrode equal the gain in weight of the other? Is this what you expected? Explain. From the atomic weight of copper (63.54 g/mole), calculate the number of moles of Cu gained or lost by each electrode. Now calculate how many moles of electrons this required (How many moles of electrons are involved in this reaction for every mole of Cu that is reduced or oxidized?).

Calculate the charge (in coulombs) that was passed through your system during this electrolysis.

At this point we know the number of electrons involved in this reaction as well as the charge passed through the system. Calculate now the number of coulombs per mole of electrons, this is the value of the Faraday, F.

Look up the charge (in C) of a single electron. Calculate the value of Avogadro's number, N. Look up the literature value of the Faraday and Avogadro's number. Calculate the % error.

Pre-Lab Questions

1. What is the value of the Faraday?
2. What is the value of the Avogadro's number?
3. What is the charge in Coulombs of a single electron?
4. If a current of 0.75 A is passed through a solution of copper(II) ions for 25 minutes. How many coulombs of charge were passed?

How many moles of copper(II) ions should have been reduced to copper metal as a result of passing this electricity through the solution?

How many grams of copper metal should have been deposited at the surface of the electrode?

