Answer on Question #77675 – Chemistry – Physical Chemistry

Why zinc does not lose any electron in water like sodium. And what happens that zinc easily loses two electrons in zinc sulphate solution in a galvanic cell.

Solution:

In chemistry, the reactivity series is a series of metals, in order of reactivity from highest to lowest. Sodium is highly active and is able to displace hydrogen from water:

$2Na\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2\ (g)$

Less active metals like zinc or iron cannot displace hydrogen from water.

The reactivity of metals is due to the difference in stability of their electron configurations as atoms and as ions. As they are all metals they will form positive ions when they react. Potassium has a single outer shell electron to lose to obtain a stable "noble gas" electron configuration. Metals that require the loss of only one electron to form stable ions are more reactive than similar metals which require the loss of more than one electron.

If you dip a single zinc electrode into some electrolyte solution, this situation is subject to this equilibrium:

$Zn(s) \leftrightarrow Zn^{2+}(aq) + 2e^-$

At first, there is no barrier to some of the zinc at the surface dissolving. The standard reduction potential is −0.7618 V, so the oxidation process is energetically favourable. However, though the zinc ion can diffuse through the solution, there is nowhere for the two electrons to go, so they are trapped within the electrode. This excess charge opposes further oxidation – it becomes more and more difficult to force more charge into the electrode. For this reason, zinc electrodes do basically nothing in neutral solutions. But if we connect the zinc electrode to another electrode that is less easily oxidized than zinc (copper is often used because it's cheap), these electrons can flow into the other electrode and participate in a reduction reaction.

A galvanic cell consists of two half-cells, each containing a metal cathode immersed in a solution of its cations, connected via a salt bridge. In our case, the two half-cells contain a zinc electrode and a copper electrode, respectively. These electrodes are immersed in copper sulfate, in the case of the copper electrode, and zinc sulfate, in the case of the zinc electrode.

Zinc is more reactive than copper, which means that it will lose electrons more readily. The redox equlibria that govern this galvanic cell are

$Zn^{2+}(aq) + 2e^- \leftrightarrow Zn(s) \qquad E^0 = -0.76\ V$

and

$Cu^{2+}(aq) + 2e^- \leftrightarrow Cu(s) \qquad E^0 = +0.34\ V$

The standard electrode potentials, E^0, listed after the two equilibrium reactions tell you which species will be oxidized, i.e. it will lose electrons, and which will be reduced, i.e. it will gain electrons. The negative E^0 for zinc tells you that the equilibrium lies to the left, meaning which confirms that zinc is more reactive and will tend to lose electrons to form Zn^{2+} cations. The positive E^0 for copper tells you that the equilibrium lies to the right, which again confirms that copper tends to hold on to its electrons and not form Cu^{2+} ions as readily.