Larger heading ?

What is the effect of the concentration of solutions and of the formation of complexes on the electromotive force of an electrochemical cell?

Background research

Redox reactions and electrode potentials

Redox reactions involve the transfer of electrons and can be split into two half equations, one producing electrons – oxidation, and the other accepting them – reduction. For example:

The formatting could be improved here I think. A new line for each reaction with some spacce above and below would be clearer.

The overall reaction is: $Cu_{(s)} + 2Ag^{+}_{(aq)} \longrightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$

and the half reactions are:

Find a larger symbol for the -

These ought to be aligned

$Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^{-}$ oxidation

$2Ag^{+}_{(aq)} + 2e^{-} \longrightarrow 2Aq_{(s)}$ reduction

Both of the individual half reactions are reversible. The direction they take is dependent on the substance with which they react [1].

If a metal strip is placed in a solution of its ions then two opposing changes can occur;

1. There is a tendency for the ions to leave the metal lattice and pass into solution.
2. There is an opposite tendency for ions from the solution to deposit on the metal.

In practice, one of these changes has a greater tendency to occur than the other, depending on the particular metal/metal ion system. Equilibrium is established between the metal and its ions in solution and a potential difference is set up between the metal strip and the solution [2]. This set-up is called a half-cell and each different half-cell has its own electrode potential (potential difference) which is determined by the position of the equilibrium [1].

Where are the references?

Could you be a little more specific here?

General comments:

Perhaps increase line spacing and indented paragraphs may be a little old fashion ?

When two half-cells are combined they form an electrochemical cell. A connection is created between the two solutions in the form of a strip of filter paper soaked in saturated potassium nitrate(v). This salt or ion bridge allows the current to be carried between the solutions via the movement of the potassium and nitate(v) ions. This ensures that, whilst there is electrical contact, there is no mixing of the solutions [1].

In one half of the electrochemical cell an oxidation reaction occurs, providing electrons for the reduction reaction that occurs in the other half of the cell. A typical set-up of an electrochemical cell is shown below:

Why should the solutions not mix?.

I don't see a diagram here, did it get lost or is that on your to-do list?

In order to compare cells by measuring the potential difference across them, a high resistance voltmeter, or potentiometer, is required to enable the measurement of the potential difference across the cell when no current is flowing, the electromotive force (e.m.f.), which is given the symbol E_{cell} or E^{θ} [1]. The electromotive force is defined as the difference in electrical potential energy per unit charge [7] and is measured in volts. It can therefore be defined as:

$$\text{EMF in volts} = \frac{\text{difference in electrical potential energy in joules}}{\text{charge in coulombs}}$$

Centre justify

This allows direct comparison because if a current were flowing in the circuit the voltage may drop. The potential difference is a measure of how much each electrode is tending to accept or release electrons.

Standard electrode potentials

The absolute potential of a single electrode or half-cell cannot be measured, but a relative potential may be determined by coupling the half-cell with a reference electrode to form a cell [3]. In this way, electrodes are assigned standard electrode potentials by coupling them with a standard hydrogen electrode, which has, by definition, an electrode potential of 0V, under standard conditions. These are taken to be when all ions taking part in the reaction are at 1.00M concentrations, all gases have a partial pressure of 1 atmosphere [7] and the temperature is at 298K.

<u>The standard hydrogen electrode</u>

The standard hydrogen electrode consists of a piece of inert platinum foil, coated electrolytically with platinum black, and immersed in a solution of hydrochloric acid containing hydrogen ions at unit activity. This corresponds to 1.8M hydrochloric acid at 25^{0}C [4]. The set-up of a standard hydrogen electrode is shown below [5]:

Another missing diagram?

Hydrogen gas under one atmosphere of pressure is passed over the platinum foil, entering through the side tube, and escapes via the small holes in the glass casing. Bubbles are periodically formed causing the level of the liquid to fluctuate. This results in part of the platinum foil being exposed to both the solution and to the hydrogen gas.

The lower end of the foil, however, is constantly immersed in the solution to avoid interruption of the electrical current. A connection is formed between the platinum foil and an external circuit using mercury.

Platinum black is used as it is able to act almost as if it is entirely composed of hydrogen. This is because it has the ability to absorb large quantities of hydrogen gas and permits the change from the gaseous to the ionic form, and the reverse, to occur without hindrance. ~~In this way it acts as a hydrogen electrode~~. Under standard conditions this system possesses a definite potential though, as mentioned earlier, it is, by definition, taken to be exactly zero volts [7] at all temperatures.

Perhaps a little too much detail.

Standard electrode potentials of some elements, E^{θ},using a hydrogen electrode as reference [6]:

Element	**E^{θ} / V**
Potassium	-2.92
Calcium	-2.87
Sodium	-2.71
Magnesium	-2.38
Aluminium	-1.66
Zinc	-0.76
Iron (Fe^{2+})	-0.44
Nickel	-0.25
Tin (Sn^{2+})	-0.14
Lead (Pb^{2+})	-0.13
Hydrogen* Copper (Cu^{2+})	0.00 +0.34
Mercury (Hg_2^{2+})	+0.79
Silver	+0.80

Electrons are lost increasingly easily

Centre justify

** defined as zero*

The order of elements given in the above table is almost the same as:
-the order found for the activity series of the metals
-the order found for the ease of discharge of cations at the cathode during electrolysis [6].

Metals usually react by losing one or more electrons to form positive ions. The ease with which the metal is able to lose these electrons is directly related to the chemical reactivity of the metal. The greater the ease with which ionisation occurs, the more reactive the metal will be. Standard electrode potentials can therefore provide a direct indication of the reactivity of the metals [6].

When two half-cells of two different electrodes are connected using a salt bridge to form an electrochemical cell, reduction will occur in one half-cell and oxidation in the other. This is dependent on the nature of the two electrodes coupled together. Electrode potentials measure the tendency of a half-reaction to accept electrons. The more positive the electrode potential, the greater is this tendency [1].

Further explanation?