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Electrochemical cells and cell potentials lab answers

The batteries in the remote controland the engine in the car are just a few examples of how chemical reactions create power through the flow of electrons. Cell aapotence is the way we can measure the voltage between the two half cells of the battery. We explain how this happens and what components allow us to find the voltage in the electrochemical cell. E_{cell} is the extent of the possible difference between the two half cells in the electrochemical cell. The possible difference is that electrons can flow from one half cell to another. Electrons are able to move between electrodes because the chemical reaction is a redox reaction. The redox reaction occurs when one substance oxidizes and another decreases. During oxidation, the material loses one or more of its electrons and thus becomes positively charged. In contrast, during the reduction, the material gains electrons and becomes negatively charged. This applies to the measurement of cell potential, since the determination of cell potential determines cell potential. Cell apothecrath (Ecell) is measured in voltage (V), which allows us to add a certain value to the cell potential. The electrochemical cell consists of two half cells. Oxidation of the metal electrode occurs in one half cell and metal ions in the solution are reduced in the other half cell. A half cell consists essentially of a metal electrode consisting of a certain metal in an aqueous solution composed of the same metal ion. The electrode is connected to the other half cell, which contains an electrode with some metal submerged in an aqueous solution of later metal ions. The first half cell, in this case, will be marked as an anode. In this half cell, the metal in the atoms in the anode oxidizes and connects to other metal ions in the aqueous solution. An example is a copper electrode in which the Cu atoms in the electrode lose two electrons and become Cu2+. Cu2+ ions are then connected to the aqueous solution, which already has a certain molarity of Cu2+ ions. Electrons lost by Cu atoms in electrodes are transferred to the second half cell, which becomes the cathode. In this example, it is assumed that the second half cell consists of a silver electrode in an aqueous solution of silver ions. As the electrons are transferred to the Ag electrode, the Ag+ ions in the solution decrease and become Ag atoms on the Ag electrode. In order to balance the filling on both sides of the cell, the half cells are connected to a salt bridge. As the anode half cell becomes overloaded with Cu2+ ions, the negative anion of salt enters the solution and stabilizes the charge. Similarly, the cathode Cell, as the solution becomes more negatively charged, cations of the salt bridge stabilize the charge. For electrons to be transferred from the anode to the cathode, there must be some kind of energy potential that makes this phenomenon favorable. The potential energy that drives redox reactions involved electrochemical cells is the potential anode oxidized and the potential for the cathode to be reduced. The electrons involved in these cells will fall into the anode, which with greater potential becomes oxidized by the cathode, which with less potential will become oxidized. This is similar to a rock that falls off a cliff, where the rock will fall from a larger potential energy to a lower potential energy. Note: The difference between the anode's reduction potential and the potential for cathode reduction is cell potential. $E^o_{Cell}= E^o_{Red,Cathode} - E^o_{Red,Anode}$ Note: Both options used in the equation are standard reduction potentials, usually found in tables such as tables P1 and P2. However, the anode reaction is actually an oxidation reaction - the reverse of the reduction reaction. That explains the minus sign. A plus sign would have been used if we had received oxidation potential $E^o_{Ox,Anode}$ since $E^o_{Red}=E^o_{Ox}$. The E^o superscript in e^o indicates that these options are correct only if the concentrations are 1 M and the pressure is 1 bar. An adjustment called the Nernst equation should be applied if the conditions are different. Here is the list of all components: Two halfcells Two metal electrodes One voltmeter One salt bridge Two aqueous solutions for each half cell Each of these components created the electrochemical cell. The voltmeter is at the very top of the gold color, which is measured by the cell voltage, or the amount of energy produced by electrodes. This value from the voltmeter is called the voltage of the electrochemical cell. It can also be called the possible difference between the half cells, Ecell. A Volt is the amount of energy for each electrical charge; 1V=1J/C: V = voltage, J=joules, C=coulomb. The voltage is basically what drives the electrons to move. If there is a high voltage, this means that there is a large movement of electrons. The voltmeter reads the transmission of electrons from the anode to the joules/coulomb cathode. The image above is called a cell chart. Cell diagram is a representation of the general reaction of the electrochemical cell. The chemicals involved are those that actually react during reduction and oxidation reactions. (Viewers will be left out). In the cell chart, the anode half cell is always written on the left side of the chart, and the cathode in the half-cell always Side. Both the anode and the cathode have two vertical lines (ll) scattered, as shown in the blue cloud above. The electrodes (yellow circles) of both anode and cathode solutions are separated by a single vertical line (l). If more than one chemical is involved in the aqueous solution, the data are added to the diagram by adding a comma and then the chemical. For example, in the image above, if copper was not oxidized alone, and another chemical like K was involved, it would denote it (Cu, K) in the figure. A cell chart makes it easier to see what oxidizes and decreases. These are the reactions that create the cell potential. Normal cell aapotence (E_{cell}) is the difference between the two electrodes that make up the cell's voltage. To find the difference between the two half cells, use the following equation: $E^o_{Cell}= E^o_{Red,Cathode} - E^o_{Red,Anode}$ E^o_{Cell} is the standard cell apotence (1 M, 1 Barr and 298 K). E^o_{Red} is the standard reduction potential of the cathode reduction half-reaction $E^o_{Red,Anode}$ is the standard reduction potential of the oxidation half-reaction in the anode. Remember, that this equation can be written as an amount, not as a difference $E^o_{Cell}= E^o_{Red,Cathode} + E^o_{Ox,Anode} \tag{1b}$ where our strategy was based on the sum of the oxidation potential and the reduction potential (which are the actual reactions) between the two reduction potentials (which are traditionally found in the reference tables). Because $E^o_{Red}=-E^o_{Ox}$, the two approaches are equivalent. The example will use an image from the Copper and Silver cell chart. The oxidation half cell of the redox equation is Cu(s) → Cu2+(aq) + 2e-EoOx= -0,340 V, where the reduction potential of EoRed= 0,340 V has been denied, which is the quantity from the list of standard reduction potentials, to find the oxidation potential of EoOx. The reduction half cell is (Ag+ + e-→ Ag(s) x2 EoRed = 0.800 V, where we multiplied the reduction chemical equation by two to balance the number of electrons, but we did not double the EoRed because the Eo values are given in voltage units. Voltage is energy per charge, not energy per reaction, so you do not have to take into account the number of reactions required to produce or consume the amount of charge used to balance the equation. Chemical equations find: Cu(s) + 2Ag+ 2e- → Cu2+(aq) + 2Ag(s) + 2e- and simplified, to find the full reaction: Cu(s) + 2Ag+ → Cu2+(aq) + 2Ag(s), where the possibilities of half-cell reactions can be summed up EoCell= EoRed,Cathode+EoOx,Anode EoCell = 0.800 V + (-0.340 V) EoCell = cell's standard cell potential is 0.460 V. We're done. Keep in mind that since $E^o_{Red}=-E^o_{Ox}$ could have achieved the same thing with the reduction potential difference, where missing or doubled negation takes into account the fact that the reverse of the reduction reaction is what actually occurs. EoCell= EoRed,Cathode-EoRed,Anode EoCell = 0.800V - 0.340V EoCell = 0.460V The table below lists the significant standard electrode potentials in the reduction state. To determine oxidation electrodes, the reduction equation can be easily reflected and its potential potential may vary from positive to negative (and vice versa). When using the following half cells, the following equation can be used instead of changing the following possible value without changing the potentials from positive to negative (and vice versa): EoCell= EoRed,Katod - EoRed,Anode Table: Reduction Half-Reaction Eo, V Acidic acid solution F2(g) + 2e-→ 2 F-(aq) +2.866 O3(g) + 2H+(aq) + 2e-→ O2(g) + H2O(l) +2.075 S2O82-(a q) + 2e-→ 2SO4 2-(aq) +2.01 H2O2(aq) + 2H+(aq) +2e-→ 2H2O(l) +1.763 MnO4-(aq) + 8H+(aq) + 5e-→ Mn2+(aq) + 4H2O(l) +1.51 PbO2(s) + 4H+(aq) + 2e-→ Pb2+(aq) + 4H2O(l) +1,455 Cl2(g) + 2e- → 2Cl-(aq) +1.1.2358 Cr2O72-(aq) + 14H+(aq) + 6e-→ 2Cr3+(aq) + 7H2O(l) +1.33 MnO2(s) + 4H+(aq) +2e- -> Mn2+(aq) + 2H2O(l) +1.23 O2(g) + 4H+(aq) + 4e-→ 2H2O(l) +1.229 2IO3-(aq) + 12H+(aq) + 10e-→ I2+(aq) + 1.20 Br2(l) + 2e-→ 2Br-(aq) +1.065 NO3-(aq) + 4H+(aq) + 3e-→ NO(g) + 2 H2O(l) +0.956 Ag+(aq) + e- → Ag+0.800 Fe3+(aq) + e-→ Fe2+(aq) +0.771 O2(g) + 2H+(ag) + 2e- → H2O2(aq) +2e 0.695 I2(s) + 2e-→ 2I-(aq) +0.535 Cu2+(aq) + 2e-→ Cu(s) +0.340 SO42-(0.535 Cu2+(aq) + 2e- → Cu(s) +0.340 SO42-(aq) + 4H+(aq) + 2e-→ H2O2(aq) +4H+(aq) + 2e-→ <9> Sn2+(aq) +0.154 S(s) + 2H+((aq) + 2e-→ H2S(g) +0.14 2H+(aq) + 2e-→ H2(g) 0 Pb+(aq) + 2e-→ Pb -0.125 Sn2+(aq) + 2e-→ Sn(s) -0.137 Fe2+(aq) + 2e-→ - Fe(s) -0.440 Zn2+ + 2e- → Zn (s) -0.763 Al3+(aq) + 3e-→ Al(s) -1. 676 Mg2+(aq) + 2e-→ Mg (s) -2,356 Na+(aq) + e-→ Na(k) -2,713 Ca2+(aq) + 2e- → Ca (k) -2.84 K+(aq) + + e- → K(s) -2,924 Li+(aq) + e-→ Li(s) -3.040 Base solution O3(aq) + H2O(l) + 2e-→ O2(g) + 2OH-((aq) +1.246 OCl-(aq) + H2O(l) + 2e-→ Cl-(aq) + 2OH-(aq) +0.890 O2(g) + 2H2O(l) +4e-→ 4OH-(aq) +0.401 2H2O(l) + + 2e-→ H2(aq) + 2OH-(aq) -0.0828 For this redox reaction $Sn+ Pb^{2+}(aq) \rightarrow Sn^{2+}(aq) + Pb (s)$ describe oxidation reduction and half-reactions. Create a chart to match the equations. In the image above, in the cell diagram, type the entire equation of the reaction. If Cu^{2+} ions in the solution are a cell cathode around a Cu metal electrode, and K^+ ions in a solution around a K metal electrode, an anode of a cell that is larger than half a cell can be reduced? What type of reaction gives the sheet of the cell How is cell potential measured and by what means is it measured? E^o_{cell} is $[4Al(s) + 3O_2(g) + 6H_2O(l) + 4OH^-(aq) \rightarrow 4[Al(OH)_4]^{-}(aq)]$ (^aq) (^) (^) (^) (aq) (^) (^) (^) equation +2.71 V. If O_2 is reduced to (OH^-) +0.401 V. What is the reduction of half-reaction? $[Al(OH)_4]^{-}(aq) + 3e^{-} \rightarrow Al(s) + 4OH^-]$ oxidation: Sn(s) → Sn2+(aq) + 2e-(aq)) reduction: Pb2+(aq) + 2e-(aq) → Pb(s) cell chart: Sn(s) cell chart: Sn(s) Sn2+(aq) || Pb2+(aq) | Pb(s) Cu(s) + 2Ag+(aq) → 2Ag +Cu2+(aq) Since the half-cell containing the Cu electrode is the cathode in solution Cu^{2+}, this is the half cell where the reduction occurs. Therefore, this half cell has greater potential to be reduced. The redox reaction. Cell potential is measured in Volt (=J/C). This can be measured using an ex-meter. You can divide the net cell equation into two half-equations. Oxidation: {Al(s) + 4OH-(aq) → [Al(OH4)]-(aq) + 3e-} x4; I don't know what happened. That's what we solve. Reduction: {O2(g) + 2H2O(l) + 4e-→ 4OH-(aq)} x3 Eo= +0.401V Net: 4Al (s) + 3O2(g) + 6H2O(l) + 4OH-(aq) → 4[Al(OH)4]-(aq) Eocell = 2. 71V Eo 2.71V= +0.401V - Eo[Al(OH)4]-(aq)Al(s)) Eo[[Al(OH)4]-(aq)}} = 0.401V - 2.71V = -2.31V Confirm this in the standard reduction potential table Petrucci, Harwood, Herring, and Madura. General chemistry: Principles and modern applications. 9. ed. Upper Saddle River, New Jersey: Pearson Education, 2007. Katherine Barrett, Gianna Navarro, Joseph Koressel, Justin Kohn Kohn

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