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# Electrochemical cells and cell potentials lab answers

MBS998 The batteries on your remote control and your car's engine are just a couple of examples of how chemical reactions create power through electron flow. Cell potential is how we can measure how much voltage exists between the two half cells in a battery. We will explain how this is done and what components allow us to find the voltage that exists in an electrochemical cell. Cell potential (E_-cell) is the measure of the potential difference between two half cells in an electrochemical cell. The potential difference is caused by the ability of electrons to flow from one half cell to the other. Electrons are able to move between electrodes because the chemical reaction is a redox reaction. One redox reaction occurs when one certain substance oxidizes, while another is reduced. During oxidation, the substance loses one or more electrons, and is therefore charged positively. In contrast, during reduction, the substance gains electrons and is charged negatively. This is related to the measurement of cellular potential because the difference between the oxidation potential of the reducing agent and the oxidizing agent will be reduced will determine the cellular potential. Cellular potential (Ecell) is measured in voltage (V), allowing us to give a certain value to cellular potential. An electrochemical cell consists of two half cells. In one half cell, oxidation of one metal electrode occurs, and in the other half of the cell, the reduction of metal ions occurs in the solution. The half cell consists essentially of a metal electrode of a given metal submerged in an aqueous solution of the same metal ions. The electrode is connected to the other half cell, which contains an electrode of some metal submerged in an aqueous solution of posterior metal ions. The first cell in the middle, in this case, will be marked as the ado. In this half cell, the metal in the atoms in the electrode oxidizes and joins the other metal ions in the aqueous solution. An example of this would be a copper electrode, in which the Cu atoms in the electrode lose two electrons and become Cu2+. The Cu2+ ions would then bind to the aqueous solution that already has a certain molarity of Cu2+ ions. The electrons lost by the Cu atoms in the electrode are transferred to the second cell, which will be the cathode. In this example, we will assume that the second cell of half consists of a silver electrode in an aqueous silver ion solution. As electrons are passed to the Ag electrode, the Ag+ ions in solution will shrink and become an Ag atom in the Ag electrode. In order to balance the load on both sides of the cell, the half cells are connected by a salt bridge. As the half cell of the aodo becomes overwhelmed with Cu2+ ions, the negative anion of salt will enter the solution and stabilize the Similarly, in the half cell of the cathode, as the solution is more negatively charged, the cations of the salt bridge the load. For electrons to be transferred from the aod to the cathode, there must be some kind of energy potential that makes this phenomenon favorable. The potential energy that drives the redox reactions involved in electrochemical cells is the potential for the anod to oxidize and the potential for the cathode to shrink. The electrons involved in these cells will fall from the anodo, which has a greater potential to oxidize to the cathode, which has less potential to oxidize. This is analogous to a rock falling from a cliff where the rock will fall from a higher potential energy to a lower potential energy. Note: The difference between the potential of the aodo to be reduced and the potential of cathode to shrink is cellular potential. Note o_ o_: Both potentials used in this equation are o_ standard reduction potentials, which are usually what is found in the tables (for example, Table P1 and Table P2). However, the reaction in the anodo is actually an oxidation reaction -- the back of a reduction reaction. This explains the minus sign. We would have used one more sign if we had been given an oxidation potential (E-o_,Ox,Anode) instead, since the o_ value of E(E-o_). The superscript or in E-o indicates that these potentials are correct only when the concentrations are 1 M and the pressures are 1 bar. A correction called Nernst Equation should be applied if the conditions are different. Here is the list of all components: Two half cells Two metal electrodes One voltmeter One salt bridge Two aqueous solutions for each half cell All these components create the electrochemical cell. The image above is an electrochemical cell. The voltmeter at the top in the golden color is what measures the cell voltage, or the amount of energy the electrodes produce. This voltmeter reading is called electrochemical cell voltage. This can also be called the potential difference between half cells, Ecell. Volts are the amount of energy for each electrical charge; 1V-1J/C: Voltage V, J-joules, C-coulomb. Voltage is basically what drives electrons to move. If there is a high voltage, that means there is a high movement of electrons. The voltmeter reads the transfer of electrons from the aodo to the cathode in Joules per Coulomb. The image above is called a cell diagram. The cell diagram is a representation of the overall reaction in the electrochemical cell. The chemicals involved are what are actually reacting during reduction and oxidation reactions. (The viewer's ions stay out).) In the cell diagram, the half cell of the aodo is always written to the left side of the diagram, and in the cathode half of the cell is always written on the right side of the diagram. Both the aodo and cathode are separated by two vertical lines (ll) and go in the blue cloud above. The electrodes (yellow circles) of acaodo and cathode solutions are by a single vertical line (l). When more chemicals are involved in the aqueous solution, they are added to the diagram by adding a comma and then the chemical. For example, in the image above, if copper does not oxidize on its own and another chemical like K was involved, it would denote it as (Cu, K) in the diagram. The cell diagram makes it easier to see what is rusting and what is shrinking. These are the reactions that create cellular potential. The difference between the two electrodes is the standard cell potential ((E-o_-cell-)) that forms the voltage of that cell. To find the difference of the two half cells, the following equation is used: E-o_, Cell (E) o_, Red, Cathode - E-o_,Red,Anode,Anode', E, E (E-o_'Cell') is the standard cell potential (less than 1M, 1 Barr and 298 K). o_ The standard reduction potential for the mean reduction reaction that occurs in the cathode (E-o_,anodo) is the standard reduction potential for the reaction of half oxidation that occurs in the anodo Units of potentials are typically measured in volts (V). Note that this equation can also be written as a sum instead of a difference. [E-o_-Cell-E-o_-Red,Cathode- + E-o_-Ox,Anode-label-1b Where we have changed our strategy of taking the difference between two reduction potentials (which are traditionally what one finds in the reference tables) to take the sum of oxidation potential and reduction potential (which are the reactions that actually occur). Since E-o_-Red-E-o_-Ox, the two approaches are equivalent. The example will use the copper and silver cell diagram image. The mean oxidation cell of the redox equation is: Cu(s) → Cu2+(aq) + 2e- EoOx -0.340 V where we have denied the reduction potential EoRed 0.340 V, which is the amount we find from a list of standard reduction potentials, to find the oxidation potential EoOx. The mean reduction cell is: ( Ag + e- → Ag(s) ) x2 EoRed to 0.800 V where we have multiplied the chemical equation of reduction by two in order to balance the electron count, but we have not duplicated EoRed since the Eo values are given in voltage units. Voltage is energy per charge, not energy per reaction, so you don't need to consider the number of reactions needed to produce or consume the amount of load you're using to balance the equation. Chemical equations can be added to find: Cu(s) + 2Ag+ + 2e- → Cu2+(aq) + 2Ag(s) + 2e- and simplified to find the general reaction: Cu(s) + 2Ag+ → Cu2+(aq) + 2Ag(s) where the potential(s)) of the half cell reactions can be added EoCell EoRed,Cathode+EoOx,Anode EoCell to 0.800 V + (-0.340 V) EoCell to 0.460V to find that the standard cell potential of this cell is 0.460 V. We are finished. Note that since E-o_-Red-E-o_-Ox, we could have achieved taking the difference in reduction potentials, when absent or duplicate denial explains the fact that the reverse reduction is what really happens. EoCell EoRed,Cathode-EoRed,Anode EoCell at 0.800V - 0.340V EoCell at 0.460V The following table is a list of important standard electrode potentials in the reduction state. To determine oxidation electrodes, the reduction equation can simply be flipped and its potential changes from positive to negative (and vice versa). When using the half cells below, instead of changing the potential of the equation below you can use it without changing any of the possibilities from positive to negative (and vice versa): EoCell EoRed,Cathode - EoRed,Anode Table: Reduction Half-Reaction Eo, V Acid solution F2(g) + 2e- → 2 F-(aq) +2.866 O3(g) + 2H+(aq) + 2e- → O2(g) + H2O(l) +2.075 S2O82-(aq ) + 2e- → 2SO42-(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((aq) + 4H2O(l) +1.455 Cl2((aq) g) + 2e- - → 2Cl-(aq) +1.358 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(s) + 6H2O(l) +1.20 Br2(l) + 2e- → 2Br-(aq) +1.065 NO3-(aq) + 4H+(aq) + 3e- → NO(g) + 2.2 H2O(l) +0.956 Ag+(aq) + e- → Ag(s) +0.800 Fe3+(aq) + e- → Fe2+(aq) +0.77 1 O2(g) + 2H+(ag) + 2e- → H2O2(aq) +0.695 I2(s) + 2e- → 2I-(aq) +0.535 Cu2+(aq) + 2e- → Cu (s) +0.340 SO42-(aq) + 4H+(aq) + 2e- → 2H2O(l) + SO2(g) +0.17 Sn4+(aq) + 2e- → Sn2+(aq) +aq) +0.154 S(s) + 2H+(aq) + 2e- → H2S(g) +0.14 2H+(aq) + 2e- → H2(g) 0 Pb2+(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) -s) 1,676 Mg2+(aq) + 2e- → Mg(s) -2,356 Na+(aq) + e- → Na(s) -2,713 Ca2+(aq) + 2e- → Ca(s) -2.84 K+(aq) + + e- → K(s) -2,924 Li+(aq) + e- → Li(s) -3,040 Basic solution O3(aq) + H2O(l) + 2e- → O2(g) + 2OH-((aq) 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 red The reaction ofox [Sn(s) + Pb-2+-(aq) s.a.a.a.-right Sn-2+-(aq) + Pb(s)-] write the mean oxidation and reduction reactions. Create a cell diagram that matches your equations. In the image above, in the cell diagram, type the general equation for the reaction. If the solution ions around a metal electrode of Cu is the cathode of a cell, and the solution ions around a metal electrode K is the acaode of a cell, which half cell has the greatest reduction potential? What kind of reaction provides the basis for cellular potential? How is cellular potential measured and with which device is it measured? The value of equation o_[4Al(s) + 3O_2(g) + 6H_2O(l) + 4OH-(aq) is +2.71 V. If the reduction of the word O_2 in the OH value) is +0.401 V. What is the mean reduction reaction for this medium reduction reaction? [[Al(OH)_4]-(aq) + 3e----rightarrow Al(s) + 4OH--] Sn(s) → Sn2+(aq) + 2e-(aq) reducción: Pb2+(aq) + 2e-(aq) → pb(s) pb(s) celdas diagram || Pb2+(aq) Pb(s) Cu(s) + 2Ag+(aq) → 2Ag(s) + Cu2+(aq) Because the half cell containing the half-coin solution electrode is the cathode, this is half the cell where the reduction is occurring. Therefore, this half cell has a greater potential to reduce. The redox reaction. Cellular potential is measured in Volts (J/C). This can be measured with the use of a voltmeter. We can divide the net cell equation into two half equations. Oxidation: Al(s) + 4OH-(aq) → [Al(OH4)]-(aq) + 3e-- x4; -Eo? This is why we're working out. 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 Eocell= 2.71V= +0.401V - Eo{Al(OH)4]-(aq)/Al(s)} Eo{[Al(OH)4]-(aq)/Al(s)} = 0.401V - 2.71V = -2.31V Confirm this on the table of standard reduction potentials References Petrucci, Harwood, Herring, and Madura. General Chemistry: Modern Principles and Applications. 9th ed. Upper Saddle River, New Jersey: Pearson Education, 2007. Katherine Barrett, Gianna Navarro, Joseph Koressel, Justin Kohn Kohn

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