I'm not a robot

When a redox operation proceeds, one species is oxidised and the other is reduced. It is useful to think of these two processes separately. For example, in the dissolution of zinc metal in acid: Oxidation: Reduction: The zinc metal is oxidized, and the protons are reduced to hydrogen gas. The splitting up of the overall process into separate parts for the oxidant and reductant is conceptual, as the individual reactions may not occur on their own. When the process is split up like this, each of the reactions for the oxidant and the reductant is known as a half-reaction. The reactant and product of each of these half-reactions are known as a redox couple, ie. H+/H2 and Zn2+/Zn. If we write the reaction for the oxidation of zinc in reverse, then it too is a reduction, and the overall process is the difference between two reduction processes. Overall process is – Reduction 1 – Reduction 2 Standard reduction potentials The overall reaction in a redox process has a characteristic Gibbs free reaction energy. If we split up the redox process into a reduction part and an oxidation part, then we can assign the overall Gibbs free reaction energy into contributions from the reduction process and the oxidation process. The overall Gibbs reaction energy is therefore the difference between the reaction energies for the two reduction half-reactions. Just as splitting up the overall redox process into two separate reduction half-reactions is conceptual, this splitting up of the reaction energy is also conceptual, and the ability to do this relies on knowing the reaction energy of one reduction half-reaction so that those of all the others can be calculated from it, as the reactions always occur in pairs. We define the free energy for the H+/H2 half-reaction as zero, and then all other reduction free energies may be calculated. In the reaction above, the overall Gibbs reaction energy may be measured, and as the H+/H2 reduction couple does not contribute, the entire reaction energy is due to the Zn2+/Zn couple. ΔG = + 147 kJ ΔG = 0 ΔG = + 147 kJ Thus, the reaction energy for the Zn2+/Zn reduction couple is known, and may be converted into a standard reduction potential, E, using the relationship ΔG = -nFE where n is the number of electrons transferred, 2 in this case, and F is Faraday's constant. So we can say the the standard reduction potential of the H+/H2 couple is zero, and that for the Zn2+/Zn couple is E = -0.76 V. The potential for the overall reaction is therefore the difference between the standard reduction potentials for the individual half-reactions. E = 0 V E = -0.76 V E = E(H+/H2) – E(Zn2+/Zn) = (0) – (-0.76) = +0.76 V The reaction as written has a positive electrode potential, and this corresponds to a negative free energy of reaction, and so the reaction is spontaneous as written. This means that the reduction of H+ ions by zinc metal is favourable under standard conditions. This further means that any reduction couple with a negative standard reduction potential, as is that for Zn2+/Zn, also has the thermodynamic tendency to reduce H+ ions. When the standard reduction potential has been calculated, we notice that those couples with negative potentials are able to reduce H+, and those with positive reduction potentials are able to oxidize H2. We can rank the couples in the order of their strength as reducing and oxidizing agents. This list is known as the electrochemical series: Ox/Red couple with strongly positive E*, where Ox is strongly oxidizing, through to Ox/Red couple with strongly negative E*, where Red is strongly reducing. Strongly Oxidizing E* = +2.87V Strongly Reducing E* = -3.04V The cell potential (voltage) for an electrochemical cell can be predicted from half-reactions and its operating conditions (chemical nature of materials, temperature, gas partial pressures, and concentrations). Determining Standard State Cell Potentials A cell's standard state potential is the potential of the cell under standard state conditions, which is approximated with concentrations of 1 mole per liter (1 M) and pressures of 1 atmosphere at 25oC. To calculate the standard cell potential for a reaction Write the oxidation and reduction half-reactions for the cell. Look up the reduction potential, Eoreduction, for the reduction half-reaction in a table of reduction potentials Look up the reduction potential for the reverse of the oxidation half-reaction and reverse the sign to obtain the oxidation potential. For the oxidation half-reaction, Eooxidation = - Eoreduction. Add the potentials of the half-cells to get the overall standard cell potential. Eocell = Eoreduction + Eooxidation Example: Find the standard cell potential for an electrochemical cell with the following cell reaction. Zn(s) + Cu2+(aq) Zn2+(aq) + Cu(s) Write the half-reactions for each process. Zn(s) Zn2+(aq) + 2 e- Cu2+(aq) + 2 e- Cu(s) Look up the standard potentials for the redcution half-reaction. Eoreduction of Cu2+ = + 0.339 V Look up the standard reduction potential for the reverse of the oxidation reaction and change the sign. Eoreduction of Zn2+ = - 0.762 V Eooxidation of Zn = - (- 0.762 V) = + 0.762 V Add the cell potentials together to get the overall standard cell potential. oxidation: Zn(s) Zn2+(aq) + 2 e- Eoox. = - Eored. = - (- 0.762 V) = + 0.762 V reduction: Cu2+(aq) + 2 e- Cu(s) Eored. = + 0.339 V overall: Zn(s) + Cu2+(aq) Zn2+(aq) + Cu(s) Eocell = + 1.101 V Top Determining Non-Standard State Cell Potentials To determine the cell potential when the conditions are other than standard state (concentrations not 1 molar and/or pressures not 1 atmosphere): Determine the standard state cell potential. Determine the new cell potential resulting from the changed conditions. Determine Q, the reaction quotient. Deternine n, the number of electrons transferred in the reaction "n". Determine Ecell, the cell potential at the non-standard state conditions using the Nernst equation. Ecell = Eocell - (RT/nF) ln Q Ecell = cell potential at non-standard state conditions Eocell = standard state cell potential R = constant (8.31 J/mole K) T = absolute temperature (Kelvin scale) F = Faraday's constant (96,485 C/mole e-) n = number of moles of electrons transferred in the balanced equation for the reaction occurring in the cell Q = reaction quotient for the reaction. aA + bB cC + dD, If the temperature of the cell remains at 25oC, the equation simplifies to: Ecell = Eocell - (0.0257/n) ln Q or in terms of log10 Ecell = Eocell - (0.0592/n) log Q Example: Predict the cell potential for the following reaction when the pressure of the oxygen gas is 2.50 atm, the hydrogen ion concentration is 0.10 M, and the bromide ion concentration is 0.25 M. O2(g) + 4 H+(aq) + 4 Br-(aq) 2 H2O(l) + 2 Br2(l) Calculate the standard cell potential for the reaction, Eocell, using the tabled values: oxidation: 4 Br-(aq) 2 Br2(l) + 4 e- Eoox. = - Eored. = - (+ 1.077 V) = - 1.077 V reduction: O2(g) + 4 H+(aq) + 4 e- 2 H2O(l) Eored. = + 1.229 V overall: O2(g) + 4 H+(aq) + 4 Br-(aq) 2 H2O(l) + 2 Br2(l) Eocell = + 0.152 V Determine the new cell potential resulting from the changed conditions. Calculate the value for the reaction quotient, Q. (Note: We calculate Q using molar concentrations for solutions and pressures for gases. Water and bromine are both liquids, therefore they are not included in the calculation of Q.) Calculate the number of moles of electrons transferred in the balanced equation, n. n = 4 moles of electrons Substitute values into the Nernst equation and solve for the non-standard cell potential, Ecell. Ecell = + 0.152 V - (0.0257/4) ln(1.02 x 106) Ecell = 0.063 V Top The Eɵ values of a species indicate how easily they can get oxidised or reducedThe more positive the value, the easier it is to reduce the species on the left of the half-equationThe reaction will tend to proceed in the forward directionThe less positive the value, the easier it is to oxidise the species on the right of the half-equationThe reaction will tend to proceed in the backward directionA reaction is feasible (likely to occur) when the Ecellɵ is positiveFor example, two half-cells in the following electrochemical cell are:Cl2 (g) + 2e- ⇌ 2Cl- (aq) Eɵ = +1.36 VCu2+ (aq) + 2e- ⇌ Cu (s) Eɵ = +0.34 VCl2 molecules are reduced as they have a more positive Eɵ valueThe chemical reaction that occurs in this half cell is:Cl2 (g) + 2e- → 2Cl- (aq) Cu2+ ions are oxidised as they have a less positive Eɵ valueThe chemical reaction that occurs in this half cell is:Cu (s) → Cu2+ (aq) + 2e-The overall equation of the electrochemical cell is (after cancelling out the electrons):Cu (s) + Cl2 (g) → 2Cl- (aq) + Cu2+ (aq)ORCu (s) + Cl2 (g) → CuCl2 (s)The forward reaction is feasible (spontaneous) as it has a positive Eɵ value of +1.02 V ((+1.36) - (+0.34))The backward reaction is not feasible (not spontaneous) as it has a negative Eɵ value of -1.02 ((+0.34) - (+1.36))A reaction is feasible when the standard cell potential Eɵ is positivePage 2The standard hydrogen electrode is a half-cell used as a reference electrode and consists of:Hydrogen gas in equilibrium with H+ ions of concentration 1.00 mol dm-3 (at 100 kPa)2H+ (aq) + 2e- ⇌ H2 (g)An inert platinum electrode that is in contact with the hydrogen gas and H+ ionsWhen the standard hydrogen electrode is connected to another half-cell, the standard electrode potential of that half-cell can be read off a high resistance voltmeterThe standard electrode potential of a half-cell can be determined by connecting it to a standard hydrogen electrodeThere are three different types of half-cells that can be connected to a standard hydrogen electrodeA metal / metal ion half-cellA non-metal / non-metal ion half-cellAn ion / ion half-cell (the ions are in different oxidation states)Example of a metal / metal ion half-cell connected to a standard hydrogen electrodeAn example of a metal/metal ion half-cell is the Ag+/ Ag half-cellAg is the metalAg+ is the metal ionThis half-cell is connected to a standard hydrogen electrode and the two half-equations are:Ag+ (aq) + e- ⇌ Ag (s) Eɵ = + 0.80 V2H+ (aq) + 2e- ⇌ H2 (g) Eɵ = 0.00 V Since the Ag+/ Ag half-cell has a more positive Eɵ value, this is the positive pole and the H+/H2 half-cell is the negative poleThe standard cell potential (Ecellɵ) is Ecellɵ = (+ 0.80) - (0.00) = + 0.80 VThe Ag+ ions are more likely to get reduced than the H+ ions as it has a greater Eɵ valueReduction occurs at the positive electrodeOxidation occurs at the negative electrodeIn a non-metal / non-metal ion half-cell, platinum wire or foil is used as an electrode to make electrical contact with the solutionLike graphite, platinum is inert and does not take part in the reactionThe redox equilibrium is established on the platinum surfaceAn example of a non-metal / non-metal ion is the Br2 / Br- half-cellBr2 is the non-metalBr- is the non-metal ionThe half-cell is connected to a standard hydrogen electrode and the two half-equations are:Br2 (aq) + 2e- ⇌ 2Br- (aq) Eɵ = +1.09 V2H+ (aq) + 2e- ⇌ H2 (g) Eɵ = 0.00 V The Br2 / Br- half-cell is the positive pole and the H+ / H2 is the negative poleThe Ecellɵ is: Ecellɵ = (+ 1.09) - (0.00) = + 1.09 VThe Br2 molecules are more likely to get reduced than H+ as they have a greater Eɵ valueExample of a non-metal / non-metal ion half-cell connected to a standard hydrogen electrodeIon / Ion half-cellA platinum electrode is again used to form a half-cell of ions that are in different oxidation statesAn example of such a half-cell is the MnO4- / Mn2+ half-cellMnO4- is an ion containing Mn with oxidation state +7The Mn2+ ion contains Mn with oxidation state +2This half-cell is connected to a standard hydrogen electrode and the two half-equations are:MnO4- (aq) + 8H+ (aq) + 5e- ⇌ Mn2+ (aq) + 4H2O (l) Eɵ = +1.52 V2H+ (aq) + 2e- ⇌ H2 (g) Eɵ = 0.00 V The H+ ions are also present in the half-cell as they are required to convert MnO4- into Mn2+ ionsThe MnO4- / Mn2+ half-cell is the positive pole and the H+ / H2 is the negative poleThe Ecellɵ is Ecellɵ = (+ 1.52) - (0.00) = + 1.52 VIons in solution half cellConventional Representation of CellsChemists use a type of shorthand convention to represent electrochemical cellsIn this convention:A solid vertical (or slanted) line shows a phase boundary, that is an interface between a solid and a solutionA double vertical line (sometimes shown as dashed vertical lines) represents a salt bridgeA salt bridge has mobile ions that complete the circuitPotassium chloride and potassium nitrate are commonly used to make the salt bridge as chlorides and nitrates are usually solubleThis should ensure that no precipitates form which can affect the equilibrium position of the half cellsThe substance with the highest oxidation state in each half cell is drawn next to the salt bridgeThe cell potential difference is shown with the polarity of the right hand electrodeThe cell convention for the zinc and copper cell would beZn (s)|Zn2+ (aq) ‖Cu2+ (aq)|Cu (s) E cell = +1.10 VThis tells us the copper half cell is more positive than the zinc half cell, so that electrons would flow from the zinc to the copperThe same cell can be written as:Cu (s)|Cu2+ (aq) ‖Zn2+ (aq)|Zn (s) E cell = -1.10 VThe polarity of the right hand half cell is negative, so we can still tell that electrons flow from the zinc to the copper half cellPage 3EquationsFind the stoichiometry for the reaction and complete the two half equations:MnO4- (aq) + 5e- 8H+ (aq) → Mn2+ (aq) + 4H2O (l)Fe2+ (aq) → Fe3+ (aq) + e- Answers: Balance the electrons:MnO4- (aq) + 5e- + 8H+ (aq) → Mn2+ (aq) + 4H2O (l)5Fe2+ (aq) → 5Fe3+ (aq) + 5e- Add the two half equations:MnO4- (aq) + 8H+ (aq) 5Fe2+ (aq) → Mn2+ (aq) + 4H2O (l) + 5Fe3+ (aq)Page 4Write complete equations for the following reactions:Hydrogen iodide reacting with sulfuric acid to form hydrogen sulfide, iodine and one other productLead(II) chloride, chlorine and one other product being formed from the reaction of concentrated hydrochloric acid with lead(IV) oxideAnswersAnswer 1:Step 1: Start with what you know:HI + H2SO4 → H2S + I2 Step 2: Consider any unaccounted for elementsThe only element that is not currently considered is oxygenStep 3: Make a common and appropriate suggestion for the missing productMost of these questions are in solution so there is always H2O, H+ and OH- availableMissing product suggestion = 4H2OHI + H2SO4 → H2S + I2 + 4H2OStep 4: Balance the remaining chemicals8HI + H2SO4 → H2S + 4I2 + 4H2OAnswer 2:Step 1: Start with what you know:PbO2 + HCl → PbCl2 + Cl2 Step 2: Consider any unaccounted elementsThe only element that is not currently considered is oxygenStep 3: Make a common and appropriate suggestion for the missing productMissing product suggestion = 2H2OPbO2 + HCl → PbCl2 + Cl2 + 2H2OStep 4: Balance the remaining chemicalsPbO2 + 4HCl → PbCl2 + Cl2 + 2H2OPage 5Exam code: H4323 hours33 questionsWhich row correctly describes oxidation and reduction in terms of the transfer of electrons and changes in oxidation state?Transfer of electronsChange in oxidation stateoxidationreductionoxidationreductionAgainlossincreasedecreaseBlossgainincreasedecreaseClossgaindecreaseincreaseDgainlossdecreaseincreaseDid this page help you? Which row describes the most common oxidation numbers of fluorine, oxygen and hydrogen in compounds?FOHA+7-1-1B+1+2+1C-7-1-1D-1-2+1Did this page help you?Below is a diagram of a voltaic cell. When the cell is running, what is happening in the salt bridge?K+ ions flow to the zinc half-cell and NO3− ions flow to the copper half-cellK+ ions flow to the copper half-cell and NO3− ions flow to the zinc half-cellK+ and NO3− ions flow to the copper half-cellK+ and NO3− ions flow to the zinc half-cellDid this page help you?Which is the correct equation to calculate electromotive force?EΘcell = EΘreduction + EΘoxidationEΘcell = EΘreduction – EΘoxidationEΘcell = EΘhydrogen + EΘreductionEΘcell = EΘoxidation – EΘreductionDid this page help you?Use the data below to help you answer this question.Half-equationEθ / VCo3+ (aq) + e– ⇌ Co2+ (aq) +1.82Cr3+ (aq) + e– ⇌ Cr2+ (aq)–0.41Fe3+ (aq) + e– ⇌ Fe2+ (aq)+0.77S2O82– (aq) + 2e– ⇌ 2SO42– (aq)+2.01Which is the most powerful oxidising agent from all the species in the table?Did this page help you?Which reaction shows hydrogen behaving as an oxidising agent?Did this page help you?21 mark1.50 g of impure iron wire was dissolved in 30 cm3 of sulfuric acid. The resulting iron(II) sulfate solution was made up to 250 cm3 with distilled water. 25.0 cm3 of the resulting solution was acidified and titrated with 0.020 mol dm-3 potassium manganate(VII) solution.20.0 cm3 of potassium manganate(VII) solution was required to reach the end-point of the titration.5Fe2+ (aq) + MnO4– (aq) + 8H+ (aq) → 5Fe3+ (aq) + Mn2+ (aq) + 4H2O (l)What is the percentage purity of the iron wire?Did this page help you?31 markA cell is constructed from the two redox systems below. Cu2+ (aq) + 2e– ⇌ Cu (s) Eθ = +0.34 V Fe2+ (aq) + 2e– ⇌ Fe (s) Eθ = -0.77 VWhich statement(s) is / are correct for the cell?The cell potential is 1.11 VThe reaction at the iron electrode is Fe (s) → Fe2+ (aq) + 2e– The copper electrode increases in massDid this page help you?The redox equilibria for a hydrogen-oxygen fuel cell in acidic conditions are shown below. 2H+ (aq) + 2e– ⇌ H2 (g) Eθ = 0.00 V ½O2 (g) + 2H+ (aq) + 2e– ⇌ H2O (l) Eθ = +1.23 VWhat is the equation for the overall cell reaction?½O2 (g) + 4H+ (aq) + 4e– → H2 (g) + H2O (l)H2 (g) + H2O (l) → ½O2 (g) + 4H+ (aq) + 4e– H2O (l) → H2 (g) + ½O2 (g)H2 (g) + ½O2 (g) → H2O (l) Did this page help you?51 markA cell is constructed from the two redox systems below. Sn2+ (aq) + 2e– ⇌ Sn (s) Eθ = -0.14 V Fe3+ (aq) + e– ⇌ Fe2+ (aq) Eθ = +0.77 VWhat is the value of the electromotive force, in V?Did this page help you?Which descriptions about the different systems used in batteries / cells are correct?Lead-acidNickel-cadmiumHydrogen-oxygenAFuelRechargeableFuelBRechargeableFuelRechargeableCFuelFuelRechargeableDRechargeableRechargeableFuelDid this page help you?The overall equation for iodine-thiosulfate titrations is shown below.2S2O32- (aq) + I2 (aq) → 2I- (aq) + S4O62- (aq)Which statement(s) is / are correct for iodine-thiosulfate titrations?Starch is added as an indicator at the start of the titrationThe thiosulfate ion acts as a reducing agentIodine-thiosulfate titrations can be used to determine the copper content in alloysDid this page help you?If a dilute acid is added to an aqueous solution containing nitrate ions, NO2-, two different nitrogen compounds are released as gases.2H+ (aq) + 2NO2- (aq) H2O (l) + NO (g) + NO2 (g)Which of the three statements below correctly describe(s) the process?1. The H+ (aq) ion is oxidised by NO2- (aq)2. Some nitrogen atoms are oxidised, and some nitrogen atoms are reduced3. The H+ (aq) ion acts as a catalystDid this page help you?A voltaic cell consisting of zinc and silver is set up. The following overall reaction takes place: Zn (s) + 2Ag+ (aq) → Zn2+ (aq) + 2Ag(s)What are the correct half-equations at each electrode?Anode (negative electrode)Cathode (positive electrode)AAg (s) → Ag+ (aq) + e-Zn2+ (aq) + 2e- → Zn (s)BAg+ (aq) + e- → Ag (s) Zn (s) → Zn2+ (aq) + 2e-CZn (s) → Zn2+ (aq) + 2e-Ag+ (aq) + e- → Ag (s) DZn2+ (aq) + 2e- → Zn (s)Ag (s) → Ag+ (aq) + e-Did this page help you?Below are four descriptions about the movements of electrons in voltaic cells.Which is the correct statement?Electrons flow through the external wire from the cathode (positive electrode) to the anode (negative electrode)Electrons flow through the external wire from the anode (negative electrode) to the cathode (positive electrode)Electrons flow through the salt bridge from the oxidizing agent to the reducing agentElectrons flow through the salt bridge from the reducing agent to the oxidizing agentDid this page help you?Some standard electrode potential data is shown in the table.Electrode half-equation EΘ / VCo3+ (aq) + e– → Co2+ (aq) +1.82Cu2+ (aq) + 2e– ⇌ Cu (s) +0.34Fe3+ (aq) + e– → Fe2+ (aq)+0.77I2 (s) + 2e– → 2I– (aq) +0.54S2O82– (aq) + 2e– → 2SO42– (aq)+2.01Which species cannot catalyse the reaction between iodide ions, I– (aq), and peroxodisulfate ions, S2O82– (aq)?Did this page help you?What happens to the hydrogen in a hydrogen-oxygen fuel cell?It is oxidised at the cathodeIt is reduced at the cathodeIt is oxidised at the anodeIt is reduced at the anodeDid this page help you?Page 6Acid-base & Redox ReactionsAcids, Acid-base Titrations, RedoxGroup 2 Elements, Group 2 CompoundsThe Halogens, Uses of Chlorine, Qualitative AnalysisProperties of Alcohols, Reactions of AlcoholsTechniques, Synthetic RoutesCarboxylic Acids & EstersCarboxylic Acids, Esters, Acyl ChloridesAmines, Preparing Amines, Amino acids, AmidesOptical Isomers, Identification of Chiral CentresCondensation Polymers, Polymer Repeat UnitsTechniques, Synthetic Routes Electrochemistry Tutorial: Galvanic Cells and the Nernst Equation >> Step 3: Calculating cell potentials Step 3: Calculating cell potentials In the preceding simulations you measured cell potentials of several cells made of various combinations of half-cells. It would be a monumental task to assemble a list of all possible cells and report a cell voltage for each. It is more useful to create a list of the potentials of half-cells and then calculate the potentials of the required combinations. Here, we will first learn how to use half-cell potentials, and then consider how to determine half-cell potentials from experiments. The voltage of an electrochemical cell depends on the redox reaction occurring in the cell. Above, we saw that it is often useful to break a redox reaction, such as: Zn(s) + Pb2+(aq) --> Zn2+(aq) + Pb(s) into the following half reactions. Half-reaction Process Half-cell location Pb2+(aq) + 2 e- --> Pb(s) Reduction Cathode (electron sink) Zn(s) --> Zn2+(aq) + 2 e- Oxidation Anode (electron source) Given how useful half-cell potentials are (you can use them to calculate the potential of any full-cell), it is not surprising that these are well tabulated in handbooks and on the web. These tables, by convention, contain the half-cell potentials for reduction. Reduction potential, the tendency to gain electrons, represents the relative tendency for a given metal ion to undergo reduction. For the above cell, the table would contain: Pb2+(aq) + 2 e- --> Pb(s) Eº = -0.13V Zn2+(aq) + 2 e---> Zn(s) Eº = -0.76V The Pb reaction in the table and in the cell are in the same direction (i.e. reduction). This will always be the case for the cathode, because the table lists reduction potentials and reduction always occurs at the cathode. However, the Zn reaction is in the opposite direction. This will always be the case for the anode, because the table lists reduction potentials and oxidation occurs at the anode. Because of this, the cell potential is the difference between the cell potential of the cathode and that of the anode: Eºcell = Eºcathode - Eºanode The only potential we can measure experimentally is that of a full cell (a reduction process must always be coupled with an oxidation process, otherwise free electrons would build up). So all measurements correspond to differences between the numbers in a table of half-cell potentials. This is analogous to the altitudes of cities. All we really care about is the difference in altitudes between cities. All tables define altitude relative to sea level, but if we instead created a table that gave the altitudes of cities relative to the top of Mt Everest (or the center of the earth), we would still get the same values for the difference in altitudes between cities. Sea level is the "reference" for altitude, meaning that sea level has an altitude of zero. For half-cell potentials, we also need to decide on a "reference". The convention among chemists is to take the following half-reaction as the standard and assign it a zero half-cell potential: 2 H+(aq) + 2 e- --> H2(g) Eº = 0.00V The electrode corresponding to this half-reaction is called the "standard hydrogen electrode". Here the notation Eº is used to represent hydrogen's standard electrode potential (declared to be zero at all temperatures). The measured cell voltage using the standard hydrogen electrode as one of the half-cells is, therefore, the potential of the other half reaction. The standard hydrogen electrode consists of hydrogen gas at 1 atmosphere and so is not convenient for us to work with in the laboratory. Instead, you will create your own table of reduction potentials using a tin half-cell as a standard instead of the hydrogen half-cell. Using the information from a standard reduction potentials table (see link below), the standard cell potential for the zinc/lead cell is then: Eºcell = Eºcathode - Eºanode = -0.13 - (- 0.76) = 0.63 V The following example shows how to use a table of standard reduction potentials [such as the one available online at: to predict the cell potentials that you measured in the simulation on the previous page: For the reaction between Cu and Sn, we can perform the following calculation: Cu2+(aq) + 2 e- --> Cu(s) Eºcathode = 0.34 (half-cell potential) Sn(s) --> Sn2+(aq) + 2 e- Eºanode = - 0.14 (half-cell potential) Since the copper half-cell is the cathode (this is where the reduction occurs) and the tin half-cell is the anode (where the oxidation occurs), our calculation would be: Eºcell = Eºcathode - Eºanode = 0.34 - (- 0.14) = 0.48 V For the reaction between Cu and Ag, we can perform the following calculation: 2 * (Ag+(aq) + 1 e- --> Ag(s)) Eºcathode = 0.80 (half-cell potential) Cu(s) --> Cu2+(aq) + 2 e- Eºanode = 0.34 (half-cell potential) Since the silver half-cell is the cathode (this is where the reduction occurs) and the copper half-cell is the anode (where the oxidation occurs), our calculation would be: Eºcell = Eºcathode - Eºanode = 0.80 - 0.34 = 0.46 V Note that we multiplied the Ag half-cell reaction by 2, to balance the electrons. However, we did not multiply the cell potential by 2. When we multiple a half-cell reaction by a constant, we do not need to alter the half-cell potential. Learning Outcomes Describe and relate the definitions of electrode and cell potentials Interpret electrode potentials in terms of relative oxidant and reductant strengths Calculate cell potentials and predict redox spontaneity using standard electrode potentials Unlike the spontaneous oxidation of copper by aqueous silver(I) ions described in section 17.2, immersing a copper wire in an aqueous solution of lead(II) ions yields no reaction. The two species, [latex]\ce{Ag+}[/latex](aq) and [latex]\ce{Pb^2+}[/latex](aq), thus show a distinct difference in their redox activity towards copper: the silver ion spontaneously oxidized copper, but the lead ion did not. Electrochemical cells permit this relative redox activity to be quantified by an easily measured property, potential. This property is more commonly called voltage when referenced in regard to electrical applications, and it is a measure of energy accompanying the transfer of charge. Potentials are measured in the volt unit, defined as one joule of energy per one coulomb of charge, V = J/C. When measured for purposes of electrochemistry, a potential reflects the driving force for a specific type of charge transfer process, namely, the transfer of electrons between redox reactants. Considering the nature of potential in this context, it is clear that the potential of a single half-cell or a single electrode can't be measured; "transfer" of electrons requires both a donor and recipient, in this case a reductant and an oxidant, respectively. Instead, a half-cell potential may only be assessed relative to that of another half-cell. It is only the difference in potential between two half-cells that may be measured, and these measured potentials are called cell potentials, Ecell, defined as [latex]\text{E}_\text{cell} = \text{E}_\text{cathode} - \text{E}_\text{anode}[/latex] where Ecathode and Eanode are the potentials of two different half-cells functioning as specified in the subscripts. As for other thermodynamic quantities, the =standard cell potential, E°cell, is a cell potential measured when both half-cells are under standard-state conditions (1 M concentrations, 1 bar pressures, 298 K): [latex]\text{E}^\circ_\text{cell} = \text{E}^\circ_\text{cathode} - \text{E}^\circ_\text{anode}[/latex] To simplify the collection and sharing of potential data for half-reactions, the scientific community has designated one particular half-cell to serve as a universal reference for cell potential measurements, assigning it a potential of exactly 0 V. This half-cell is the standard hydrogen electrode (SHE) and it is based on half-reaction below: [latex]\ce{2H+} (aq) +2e^- \rightarrow \ce{H2} (g)[/latex] A typical SHE contains an inert platinum electrode immersed in precisely 1 M aqueous [latex]\ce{H+}[/latex] and a stream of bubbling [latex]\ce{H2}[/latex] gas at 1 bar pressure, all maintained at a temperature of 298 K (see Figure 17.3.1). Figure 17.3.1. A standard hydrogen electrode (SHE). The assigned potential of the SHE permits the definition of a conveniently measured potential for a single half-cell. The electrode potential (EX) for a half-cell X is defined as the potential measured for a cell comprised of X acting as cathode and the SHE acting as anode: [latex]\text{E}_\text{cell} = \text{E}_X - \text{E}_\text{SHE}[/latex] [latex]\text{E}_\text{SHE} = 0 \text{V} (\text{defined})[/latex] [latex]\text{E}_\text{cell} = \text{E}_X[/latex] When the half-cell X is under standard-state conditions, its potential is the standard electrode potential, E°X. Since the definition of cell potential requires the half-cells function as cathodes, these potentials are sometimes called standard reduction potentials. This approach to measuring electrode potentials is illustrated in Figure 17.3.2, which depicts a cell comprised of an SHE connected to a copper(II)/copper(0) half-cell under standard-state conditions. A voltmeter in the external circuit allows measurement of the potential difference between the two half-cells. Since the [latex]\ce{Cu}[/latex] half-cell is designated as the cathode in the definition of cell potential, it is connected to the red (positive) input of the voltmeter, while the designated SHE anode is connected to the black (negative) input. These connections insure that the sign of the measured potential will be consistent with the sign conventions of electrochemistry per the various definitions discussed above. A cell potential of +0.337 V is measured, and so [latex]\text{E}^\circ_\text{cell} = \text{E}^\circ_\text{Cu} = +0.377 \text{ V}[/latex] Tabulations of E° values for other half-cells measured in a similar fashion are available as reference literature to permit calculations of cell potentials and the prediction of the spontaneity of redox processes. Figure 17.3.2. A cell permitting experimental measurement of the standard electrode potential for the half-reaction [latex]\ce{Cu^2+}(aq)+2e- \longrightarrow \ce{Cu}(s)[/latex] Table 17.3.1 provides a listing of standard electrode potentials for a selection of half-reactions in numerical order, and a more extensive listing is given in Standard Electrode (Half-Cell) Potentials. Table 17.3.1. Selected Standard Reduction Potentials at 25 °C Half-Reaction E° (V) [latex]\ce{F2}(g)+\ce{2e-}\longrightarrow \ce{2F-}(aq)[/latex] +2.866 [latex]\ce{PbO2}(s)+\ce{SO4^2-}(aq)+\ce{4H+}(aq)+\ce{2e-}\longrightarrow \ce{PbSO4}(s)+\ce{2H2O}(l)[/latex] +1.69 [latex]\ce{MnO4-}(aq)+\ce{8H+}(aq)+\ce{5e-}\longrightarrow \ce{Mn^2+}(aq)+\ce{4H2O}(l)[/latex] +1.507 [latex]\ce{Au^3+}(aq)+\ce{3e-}\longrightarrow \ce{Au}(s)[/latex] +1.498 [latex]\ce{Cl2}(g)+\ce{2e-}\longrightarrow \ce{2Cl-}(aq)[/latex] +1.35827 [latex]\ce{O2}(g)+\ce{4H+}(aq)+\ce{4e-}\longrightarrow \ce{2H2O}(l)[/latex] +1.229 [latex]\ce{Pt^2+}(aq)+\ce{2e-}\longrightarrow \ce{Pt}(s)[/latex] +1.20 [latex]\ce{Br2}(aq)+\ce{2e-}\longrightarrow \ce{2Br-}(aq)[/latex] +1.0873 [latex]\ce{Ag+}(aq)+\ce{e-}\longrightarrow \ce{Ag}(s)[/latex] +0.7996 [latex]\ce{Hg2^2+}(aq)+\ce{2e-}\longrightarrow \ce{2Hg}(l)[/latex] +0.7973 [latex]\ce{Fe^3+}(aq)+\ce{e-}\longrightarrow \ce{Fe^2+}(aq)[/latex] +0.771 [latex]\ce{MnO4-}(aq)+\ce{2H2O}(l)+\ce{3e-}\longrightarrow \ce{MnO2}(s)+\ce{4OH-}(aq)[/latex] +0.558 [latex]\ce{I2}(s)+\ce{2e-}\longrightarrow \ce{2I-}(aq)[/latex] +0.5355 [latex]\ce{NiO2}(s)+\ce{2H2O}(l)+\ce{2e-}\longrightarrow \ce{Ni(OH)2}(s)+\ce{2OH-}(aq)[/latex] +0.49 [latex]\ce{Cu^2+}(aq)+\ce{2e-}\longrightarrow \ce{Cu}(s)[/latex] +0.337 [latex]\ce{Hg2Cl2}(s)+\ce{2e-}\longrightarrow \ce{2Hg}(l)+\ce{2Cl-}(aq)[/latex] +0.26808 [latex]\ce{AgCl}(s)+\ce{e-}\longrightarrow \ce{Ag}(s)+\ce{Cl-}(aq)[/latex] +0.22233 [latex]\ce{Sn^4+}(aq)+\ce{2e-}\longrightarrow \ce{Sn^2+}(aq)[/latex] +0.151 [latex]\ce{2H+}(aq)+\ce{2e-}\longrightarrow \ce{H2}(g)[/latex] 0.00 [latex]\ce{Pb^2+}(aq)+\ce{2e-}\longrightarrow \ce{Pb}(s)[/latex] −0.126 [latex]\ce{Sn^2+}(aq)+\ce{2e-}\longrightarrow \ce{Sn}(s)[/latex] −0.1262 [latex]\ce{Ni^2+}(aq)+\ce{2e-}\longrightarrow \ce{Ni}(s)[/latex] −0.257 [latex]\ce{Co^2+}(aq)+\ce{2e-}\longrightarrow \ce{Co}(s)[/latex] −0.28 [latex]\ce{PbSO4}(s)+\ce{2e-}\longrightarrow \ce{Pb}(s)+\ce{SO4^2-}(aq)[/latex] −0.3505 [latex]\ce{Cd^2+}(aq)+\ce{2e-}\longrightarrow \ce{Cd}(s)[/latex] −0.4030 [latex]\ce{Fe^2+}(aq)+\ce{2e-}\longrightarrow \ce{Fe}(s)[/latex] −0.447 [latex]\ce{Cr^3+}(aq)+\ce{3e-}\longrightarrow \ce{Cr}(s)[/latex] −0.744 [latex]\ce{Mn^2+}(aq)+\ce{2e-}\longrightarrow \ce{Mn}(s)[/latex] −1.185 [latex]\ce{Zn(OH)2}(s)+\ce{2e-}\longrightarrow \ce{Zn}(s)+\ce{2OH-}(aq)[/latex] −1.245 [latex]\ce{Zn^2+}(aq)+\ce{2e-}\longrightarrow \ce{Zn}(s)[/latex] −0.7618 [latex]\ce{Al^3+}(aq)+\ce{3e-}\longrightarrow \ce{Al}(s)[/latex] −1.662 [latex]\ce{Mg^2}(aq)+\ce{2e-}\longrightarrow \ce{Mg}(s)[/latex] −2.372 [latex]\ce{Na+}(aq)+\ce{e-}\longrightarrow \ce{Na}(s)[/latex] −2.71 [latex]\ce{Ca^2+}(aq)+\ce{2e-}\longrightarrow \ce{Ca}(s)[/latex] −2.868 [latex]\ce{Ba^2+}(aq)+\ce{2e-}\longrightarrow \ce{Ba}(s)[/latex] −2.912 [latex]\ce{K+}(aq)+\ce{e-}\longrightarrow \ce{K}(s)[/latex] −2.931 [latex]\ce{Li+}(aq)+\ce{e-}\longrightarrow \ce{Li}(s)[/latex] −3.04 What is the standard potential of the galvanic cell shown in Figure 17.2.2? [reveal-answer q="576838"]Show Solution[/reveal-answer] [hidden-answer a="576838"] The cell in Figure 17.2.2 is galvanic, the spontaneous cell reaction involving oxidation of its copper anode and reduction of silver(I) ions at its silver cathode: [latex]\begin{array}{rl}{}\ce{cell reaction:}& \ce{Cu}(s) + \ce{2Ag+} (aq) \longrightarrow \ce{Cu^2+} (aq) + \ce{2Ag} (s)\\ \text{anode half-reaction:}& \ce{Cu}(s) \longrightarrow \ce{Cu^2+} (aq) + \ce{2e-}\\\text{cathode half-reaction:}& \ce{2Ag+} (aq) + \ce{2e-} \longrightarrow \ce{2Ag} (s)\end{array}[/latex] The standard cell potential computed as [latex]\begin{array}{rl}{}\text{E}^{\circ}_{\text{cell}} &=& \text{E}^{\circ}_{\text{cathode}} - \text{E}^{\circ}_{\text{anode}}\\ &=& \text{E}^{\circ}_{\ce{Ag}} - \text{E}^{\circ}_{\ce{Cu}}\\ &=& 0.7996 \text{V} - 0.34 \text{V}\\ &=&+0.46 \text{V}\end{array}[/latex] [/hidden-answer] Check Your Learning Intrepreting Electrode and Cell Potentials Thinking carefully about the definitions of cell and electrode potentials and the observations of spontaneous redox change presented thus far, a significant relation is noted. The previous section described the spontaneous oxidation of copper by aqueous silver(I) ions, but no observed reaction with aqueous lead(II) ions. Results of the calculations in Example 1 have just shown the spontaneous process is described by a positive cell potential while the nonspontaneous process exhibits a negative cell potential. And so, with regard to the relative effectiveness ("strength") with which aqueous [latex]\ce{Ag+}[/latex] and [latex]\ce{Pb^2+}[/latex] ions oxidize [latex]\ce{Cu}[/latex] under standard conditions, the stronger oxidant is the one exhibiting the greater standard electrode potential, E°. Since by convention electrode potentials are for reduction processes, an increased value of E° corresponds to an increased driving force behind the reduction of the species (hence increased effectiveness of its action as an oxidizing agent on some other species). Negative values for electrode potentials are simply a consequence of assigning a value of 0 V to the SHE, indicating the reactant of the half-reaction is a weaker oxidant than aqueous hydrogen ions. Applying this logic to the numerically ordered listing of standard electrode potentials in Table 1 shows this listing to be likewise in order of the oxidizing strength of the half-reaction's reactant species, decreasing from strongest oxidant (most positive E°) to weakest oxidant (most negative E°). Predictions regarding the spontaneity of redox reactions under standard state conditions can then be easily made by simply comparing the relative positions of their table entries. By definition, E°cell is positive when E°cathode > E°anode, and so any redox reaction in which the oxidant's entry is above the reductant's entry is predicted to be spontaneous. Reconsideration of the two redox reactions in Example 1 provides support for this fact. The entry for the silver(I)/silver(0) half-reaction is above that for the copper(II)/copper(0) half-reaction, and so the oxidation of [latex]\ce{Cu}[/latex] by [latex]\ce{Ag+}[/latex] is predicted to be spontaneous (E°cathode > E°anode and so E°cell > 0). Conversely, the entry for the lead(II)/lead(0) half-cell is beneath that for copper(II)/copper(0), and the oxidation of [latex]\ce{Cu}[/latex] by [latex]\ce{Pb^2+}[/latex] is nonspontaneous (E°cathode < E°anode and so E°cell < 0). Recalling the chapter on thermodynamics, the spontaneities of the forward and reverse reactions of a reversible process show a reciprocal relationship: if a process is spontaneous in one direction, it is non-spontaneous in the opposite direction. As an indicator of spontaneity for redox reactions, the potential of a cell reaction shows a consequential relationship in its arithmetic sign. The spontaneous oxidation of copper by lead(II) ions is not observed, [latex]\ce{Cu} (s) +\ce{Pb^2+} (aq) \rightarrow \ce{Cu^2+} (aq) +\ce{Pb} (s)[/latex] [latex]E^\circ _{forward} = -0.47 V (\text{negative, non-spontaneous})[/latex] and so the reverse reaction, the oxidation of lead by copper(II) ions, is predicted to occur spontaneously: [latex]\ce{Pb} (s) + \ce{Cu^2+} (aq) \rightarrow \ce{Pb^2+} (aq) + \ce{Cu} (s)[/latex] [latex]E^\circ _{forward} = +0.47 V (\text{positive, spontaneous})[/latex] Note that reversing the direction of a redox reaction effectively interchanges the identities of the cathode and anode half-reactions, and so the cell potential is calculated from electrode potentials in the reverse subtraction order than that for the forward reaction. In practice, a voltmeter would report a potential of −0.47 V with its red and black inputs connected to the [latex]\ce{Pb}[/latex] and [latex]\ce{Cu}[/latex] electrodes, respectively. If the inputs were swapped, the reported voltage would be +0.47 V. Are aqueous iron(II) ions predicted to spontaneously oxidize elemental chromium under standard state conditions? Assume the half-reactions to be those available in Table 17.3.1. [reveal-answer q="576840"]Show Solution[/reveal-answer] [hidden-answer a="576840"] Referring to the tabulated half-reactions, the redox reaction in question can be represented by the equations below: [latex]\ce{Cr} (s) + \ce{Fe^2+} (aq) \longrightarrow \ce{Cr^3+} (aq) + \ce{Fe} (s)[/latex] The entry for the putative oxidant, [latex]\ce{Fe^2+}[/latex], appears above the entry for the reductant, [latex]\ce{Cr}[/latex], and so a spontaneous reaction is predicted per the quick approach described above. Supporting this predication by calculating the standard cell potential for this reaction gives [latex]\begin{array}{rl}{}\text{E}^{\circ}_{\text{cell}} &=& \text{E}^{\circ}_{\text{cathode}} - \text{E}^{\circ}_{\text{anode}}\\ &=& \text{E}^{\circ}_{\text{Fe(II)}} - \text{E}^{\circ}_{\text{Cr}}\\ &=& -0.447 \text{V} - (-0.774) \text{V}\\ &=&+0.297 \text{V}\end{array}[/latex] The positive value for the standard cell potential indicates the process is spontaneous under standard state conditions. [/hidden-answer] Check Your Learning The property of potential, E, is the energy associated with the separation/transfer of charge. In electrochemistry, the potentials of cells and half-cells are thermodynamic quantities that reflect the driving force or the spontaneity of their redox processes. The cell potential of an electrochemical cell is the difference in between its cathode and anode. To permit easy sharing of half-cell potential data, the standard hydrogen electrode (SHE) is assigned a potential of exactly 0 V and used to define a single electrode potential for any given half-cell. The electrode potential of a half-cell, EX, is the cell potential of said half-cell acting as a cathode when connected to a SHE acting as an anode. When the half-cell is operating under standard state conditions, its potential is the standard electrode potential, E°X. Standard electrode potentials reflect the relative oxidizing strength of the half-reaction's reactant, with stronger oxidants exhibiting larger (more positive) E°X values. Tabulations of standard electrode potentials may be used to compute standard cell potentials, E°cell, for many redox reactions. The arithmetic sign of a cell potential indicates the spontaneity of the cell reaction, with positive values for spontaneous reactions and negative values for nonspontaneous reactions (spontaneous in the reverse direction). Key Equations [latex]{E}_{\text{cell}}^{\circ }={E}_{\text{cathode}}^{\circ }-{E}_{\text{anode}}^{\circ }[/latex] For each reaction listed, determine its standard cell potential at 25 °C and whether the reaction is spontaneous at standard conditions. [latex]\ce{Mg}(s)+\ce{Ni^2+}(aq)\longrightarrow \ce{Mg^2+}(aq)+\ce{Ni}(s)[/latex] [latex]\ce{2Ag+}(aq)+\ce{Cu}(s)\longrightarrow \ce{Cu^2+}(aq)+\ce{2Ag}(s)[/latex] [latex]\ce{Mn}(s)+\ce{Sn(NO3)2}(aq)\longrightarrow \ce{Mn(NO3)2}(aq)+\ce{Sn}(s)[/latex] [latex]\ce{3Fe(NO3)2}(aq)+\ce{Au(NO3)3}(aq) \longrightarrow \ce{3Fe(NO3)3}(aq)+\ce{Au}(s)[/latex] For each reaction listed, determine its standard cell potential at 25 °C and whether the reaction is spontaneous at standard conditions. [latex]\ce{Mn}(s)+\ce{Ni^2+}(aq)\longrightarrow \ce{Mn^2+}(aq)+\ce{Ni}(s)[/latex] [latex]\ce{3Cu^2+}(aq)+\ce{2Al}(s)\longrightarrow \ce{2Al^3+}(aq)+\ce{2Cu}(s)[/latex] [latex]\ce{Na}(s)+\ce{LiNO3}(aq)\longrightarrow \ce{NaNO3}(aq)+\ce{Li}(s)[/latex] [latex]\ce{Ca(NO3)2}(aq)+\ce{Ba}(s)\longrightarrow \ce{Ba(NO3)2}(aq)+\ce{Ca}(s)[/latex] Determine the overall reaction and its standard cell potential at 25 °C for this reaction. Is the reaction spontaneous at standard conditions? [latex]\ce{Cu}(s)\mid \ce{Cu^2+}(aq)\parallel \ce{Au^3+}(aq)\mid \ce{Au}(s)[/latex] Determine the overall reaction and its standard cell potential at 25 °C for the reaction involving the galvanic cell made from a half-cell consisting of a silver electrode in 1 M silver nitrate solution and a half-cell consisting of a zinc electrode in 1 M zinc nitrate. Is the reaction spontaneous at standard conditions? Determine the overall reaction and its standard cell potential at 25 °C for the reaction involving the galvanic cell in which cadmium metal is oxidized to 1 M cadmium(II) ion and a half-cell consisting of an aluminum electrode in 1 M aluminum nitrate solution. Is the reaction spontaneous at standard conditions? Determine the overall reaction and its standard cell potential at 25 °C for these reactions. Is the reaction spontaneous at standard conditions? Assume the standard reduction for [latex]\ce{Br2}[/latex](l) is the same as for [latex]\ce{Br2}[/latex](aq).[latex]\ce{Pt}(s)\mid \ce{H2}(g)\mid \ce{H+}(aq)\parallel \ce{Br2}(aq)\mid \ce{Br-}(aq)\mid \ce{Pt}(s)[/latex] [reveal-answer q="802553"]Show Selected Solutions[/reveal-answer] [hidden-answer a="802553"] 1. The answers are as follows: [latex]{E}_{\text{cell}}^{\circ }={E}_{\text{cathode}}^{\circ }-{E}_{\text{anode}}^{\circ }=-\text{0.257 V}-\left(-\text{2.372 V}\right)=\text{+2.115 V (spontaneous)}[/latex] [latex]{E}_{\text{cell}}^{\circ }={E}_{\text{cathode}}^{\circ }-{E}_{\text{anode}}^{\circ }=\text{0.7996 V}-\left(\text{+0.337 V}\right)=\text{+0.4626 V (spontaneous)}[/latex] [latex]{E}_{\text{cell}}^{\circ }={E}_{\text{cathode}}^{\circ }-{E}_{\text{anode}}^{\circ }=-\text{0.1262 V}-\left(-\text{1.185 V}\right)=\text{+1.0589 V (spontaneous)}[/latex] [latex]{E}_{\text{cell}}^{\circ }={E}_{\text{cathode}}^{\circ }-{E}_{\text{anode}}^{\circ }=\text{1.498 V}-\left(\text{+0.771 V}\right)=\text{+0.727 V (spontaneous)}[/latex] 3. The reaction occurs as follows: [latex]\begin{array}{rl}\text{anode:}&3\times \left(\ce{Cu}\left(s\right)\longrightarrow {\ce{Cu}}^{2+}\left(aq\right)+{\text{2e}}^{-}\right){E}_{{\ce{Cu}}^{2+}\ce{/Cu}}^{\circ }\\ \text{cathode:}&2\times \left({\ce{Au}}^{3+}\left(aq\right)+{\text{3e}}^{-}\longrightarrow \ce{Au}\left(s\right)\right)\\ \\ \text{overall:}&^3\ce{Cu}\left(s\right)+{\ce{2Au}}^{3+}\left(aq\right)\longrightarrow {\ce{3Cu}}^{2+}\left(aq\right)+\ce{2Au}\left(s\right)\end{array}[/latex] [latex]{E}_{\text{cell}}^{\circ }={E}_{\text{cathode}}^{\circ }-{E}_{\text{anode}}^{\circ }=1.4\text{98 V}-\left(\text{+0.34 V}\right)=\text{+1.16 V}\left(\text{spontaneous}\right)[/latex] 5. Oxidation occurs at the anode and reduction at the cathode: [latex]\begin{array}{rl}{}\text{anode:}&3\times \left(\ce{Cd}\left(s\right)\longrightarrow {\ce{Cd}}^{2+}\left(aq\right)+{\text{2e}}^{-}\right){E}_{{\text{Cd}}^{2+}\text{/Cd}}^{\circ }=-\text{0.4030 V}\\ \text{cathode:}&2\times \left({\ce{Al}}^{3+}\left(aq\right)+{\text{3e}}^{-}\longrightarrow \ce{Al}\left(s\right)\right) {E}_{{\ce{Al}}^{3+}\ce{/Al}}^{\circ }=-\text{1.662 V}\\ \\ \text{overall:}&3\ce{Cd}\left(s\right)+{\ce{2Al}}^{3+}\left(aq\right)\longrightarrow {\text{3Cd}}^{2+}\left(aq\right)+\ce{2Al}\left(s\right)\end{array}[/latex] [latex]{E}_{\text{cell}}^{\circ }={E}_{\text{cathode}}^{\circ }-{E}_{\text{anode}}^{\circ }=-\text{1.662 V}-\left(-\text{0.4030 V}\right)=-\text{1.259 V}\left(\text{nonspontaneous}\right)[/latex] [/hidden-answer] Glossary standard cell potential [latex]\left({E}_{\text{cell}}^{\circ }\right)[/latex]: the cell potential when all reactants and products are in their standard states (1 bar or 1 atm or gases; 1 M for solutes), usually at 298.15 K; can be calculated by subtracting the standard reduction potential for the half-reaction at the anode from the standard reduction potential for the half-reaction occurring at the cathode standard hydrogen electrode (SHE): the electrode consists of hydrogen gas bubbling through hydrochloric acid over an inert platinum electrode whose reduction at standard conditions is assigned a value of 0 V; the reference point for standard reduction potentials standard reduction potential (E°): the value of the reduction under standard conditions (1 bar or 1 atm for gases; 1 M for solutes) usually at 298.15 K; tabulated values used to calculate standard cell potentials Licenses and Attributions (Click to expand) CC licensed content, Shared previously Chemistry 2e. 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License Terms: Standard YouTube License definitionthe cell potential when all reactants and products are in their standard states (1 bar or 1 atm or gases; 1 M for solutes), usually at 298.15 K; can be calculated by subtracting the standard reduction potential for the half-reaction at the anode from the standard reduction potential for the half-reaction occurring at the cathode the electrode consists of hydrogen gas bubbling through hydrochloric acid over an inert platinum electrode whose reduction at standard conditions is assigned a value of 0 V; the reference point for standard reduction potentials the value of the reduction under standard conditions (1 bar or 1 atm for gases; 1 M for solutes) usually at 298.15 K; tabulated values used to calculate standard cell potentials