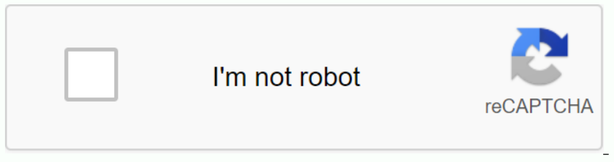

Continue

Writing half reactions worksheet answers

Determine the oxidation number of the elements in each of the following compounds: a. H2CO3 H: +1, O: -2, C: +4 b. N2 N: 0 c. Zn(OH)42- Zn: 2+, H: +1, O: -2 d. NO2- N: +3, O: -2 e. LiH Li: +1, H: -1 f. Fe3O4 Faith: +8/3, O: -2 Identify the species that oxidize and reduce in each of the following reactions: a. Cr+ + Sn4+ Cr3+ + Sn2+ Cr+: oxidized, Sn4+: reduced b. 3 Hg2+ + 2 Fe(s) 3 Hg2 + 2 Fe3+ Hg2+: reduced, Faith: oxidized c. 2 Ace(s) + 3 Cl2 (g) 2 AsCl3 As: oxidized, Cl2: reduced Would you use an oxidizing agent or reducing agent to cause the following reactions? a. ClO3- ClO2 reducing agent b. SO42- S2- reducing agent c. Mn2+ MnO2 oxidizing agent d. Oxidizing agent ZnCl2 Write balanced equations for the following redox reactions: a. 2 NaBr + Cl2 2 NaCl + Br2 b. Fe2O3 + 3 CO 2 Faith + 3 CO2 in acid solution c. 5 CO + I2O5 5 CO2 + I2 in basic solution Write balanced equations for the following reactions: a. Cr(OH)3 + Br2 CrO42- + Br- in basic solution 10 OH- + 2 Cr(OH)3 + 3 Br2 2 CrO42- + 8 H2O + 6 Br- b. O2 + Sb H2O2 + SbO2- in basic solution 2 OH- + 2 Sb + 3 O2 + 2 H2O 2 SbO2- + 3 H2O2 c. HCOOH + MnO4- CO2 + Mn2+ in acid solution 6 H+ + 2 MnO4- + 5 HCOOH 2 Mn2+ + 8 H2O + 5 CO2 d2. ClO2- ClO2 + Cl- in acid solution 5 ClO2- + 4 H+ 4 ClO2 + Cl- + 2 H2O Write the balanced mean reactions of the following reactions: a. NiO2 + 2 H2O + Fe Ni(OH)2 + Fe(OH)2 in basic solution 2 H2O + NiO2 + 2 e- Ni(OH)2 + 2 OH- 2 OH- + Fe(OH)2 + 2 e- b. CO2 + 2 NH2OH CO + N2 + 3 H2O in basic solution CO2 + H2O + 2 e- CO + 2 OH- 2 OH- + 2 NH2OH N2 + 2 e- + 4 H2O c. 2 H+ + H2O2 + 2 Fe2+ 2 Fe3+ + 2 H2O in acid solution H2O2 + 2 e- + 2 H+ 2 H2 OR Fe2+ Fe3+ + e- d. H+ + 2 H2O + 2 MnO4- + 5 SO2 2 Mn2+ + 5 HSO4- in acid solution 8 H+ + MnO4- + 5 e- Mn2+ + 4 H2O SO2 + 2 H2O HSO4- + 3 H+ + 2 e- Oxidation and reduction questions Feeding point lesson- double click on the lesson number. Quiz 1 worksheets. Oxidation, Reduction, Agents and Reactions. WS 1 2. Laboratory: The strength of oxidizing agents. Laboratory Brochure 3. Oxidation Numbers Spontaneous reactions WS 2 4. Oxidation numbers, application to reactions. WS 3 1 5. Balance Redox Half Acid/Base. WS 4 6. Balance of Redox reactions in acid/base. WS 5 2 7. Standard potentials using the chart. WS 6 3 8. Electrochemical cells. WS 7 9. Electrochemical Cell Laboratory. 10. Electrolyte cells. WS 8 4 11. Electrolytic Cell Laboratory. 12. Application of Electrolytic Cells. WS 9 5 13. Application of Electrochemical Cells: Bat & Cor. WS 10 6 14. Breathalyzer and review. Practice test WS 11 no 1 15. Review WS No. 12 Practice Test No. 2 16. Revision 17. Test. Textbook Hebden Read Unit V If you want an A in this class you need to do this!! Redox Half Reactions Reactions and Reactions WS #1 Define Every 1. Oxidation - loss of electrons 2. Reduction - electron gain 3. Oxidizing agent - causes oxidation when subjected to a reduction 4. Reducing agent - causes reduction when undergoing oxidation Write half reactions for each of the following atoms or ions. Label each as oxidation or reduction. 5. Al Al3+ + 3e- oxidation 6. S + 2e- ---------> S2- reduction 7. 2O2- ---------- > O2 + 4e- oxidation 8. Ba2+ + 2e- ----------> Ba reduction 9. 2N3- ---------- > N2 + 6e- oxidation 10. Br2 + 2e- --------- > 2Br- reduction 11. P + 3e- ---------- > P3- reduction 12. Ca ---------- > Ca2+ + 2e- oxidation 13 Ga3+ + 3e- ----------> Ga reduction 14. S + 2e- ---------> S2- reduction 15. H2 --------- > 2H+ + 2e- oxidation 16. 2H+ + 2e- ---------> H2 reduction 17. 2F- ----------> F2 + 2e- 18. P3- ---------- > P + 3e- oxidation Balance each spontaneous redox equation. Identify reduced and oxidized features. Indicate the reducing agent and oxidizing agent. 19. Al & Zn2+ 2Al + 3Zn2+ → 2Al3+ + 3Zn Oxidized Oxidizing Agent Reducing Agent 20. F2 & O2- 2F2 + 2O2- → 4F- + O2 oxidizing agent oxidized reduced reducing agent 21. O2 & Ca 2Ca + O2 → 2Ca2+ + 2O2- oxidized reducing agent oxidizing agent 22. Al3+ & Li Al3+ + 3Li → Al + 3Li+ Oxidizing Agent Oxidized Reduced Reducing Agent Reducer Agent Labels the reduced species, which oxidizes, the reducing agent and the oxidizing agent. 23. Fe2+ + Co Co2+ + Fe Co → Co2+ + 2e- oxidation Fe2+ + 2e- → Fe reduction 24. 3 Ag+ + Ni → Ni3+ + 3 Ag Ni → Ni3+ + 3e- oxidation Ag+ + 1e- → Ag reduction 25. Cu2+ + Pb → Pb2+ + Cu Pb → Pb2+ + 2e- oxidation Cu2+ + 2e- → Cu reduction 26. O2 + 2 Sn → O2- + 2 Sn2+ Sn → Sn2+ + 2e- oxidation O2 + 4e- → 2O2- reduction 27. Co2+ + 2 F- → Co + F2 2F- → F2 + 2e- oxidation Co2+ + 2e- → Co reduction 28. List the species (formulas above) that lose electrons: Co Ni Pb Sn F- 29. List the species (formulas above) that gain electrons: Fe2+ Ag+ Cu2+ O2 Co2+ For each of the following reactions, identify: -The oxidizing agent. -Agent Reducer. -The oxidized substance. -The reduced substance. 30. I- + Cl2 ---------- > Cl- + I2 Oxidized substance I- Reducing agent I- Oxidizing agent Cl2 Reduced substance Cl2 31. Co + Fe3+ ----------> Co2+ + Fe2+ Oxidized Substance Co Reducing Agent Co Oxidizing Agent Fe3+ Reduced Substance Fe3+ 32. Cr6+ + Fe2+ ----------> Cr3+ + Fe3+ Oxidized Substance Fe2+ Reducing Agent Fe2+ Oxidizing Agent Cr6+ Reduced Substance Cr6+ Redox Media Reactions reactions and reactions WS #2 1. Enter the oxidation number of each of the underlined elements. a) NH3 -3 b) H2SO4 6 c) ZnSO3 4 d) Al(OH)3 3 e) Na 0 f) Cl2 0 g) AgNO3 5 h) Cl O4- 7 i) SO2 4 j) K2Cr2O4 3 k) Ca(ClO3)2 5 l) K2Cr2O7 6 m) HPO32- 3 n) 1 o) MnO2 4 p) KClO3 5 q) PbO2 4 r) PbSO4 2 s) K2SO4 6 t) NH4+ -3 u) Na2O2 -1 v) FeO 2 w) Fe2O3 3 x) SiO44- -2 y) NaIO 3 z) ClO3- 5 aa) NO3- 5 bb) Cr(OH)4 4 cc) CaH2 -1 dd) Pt(H20)5(OH)2+ +3 ee) Fe(H2O)63+ +3 ff) CH3COOH 0 2. What is the number of carbon oxidation in each of the following substances? a) CO 2 b) C 0 c) CO2 4 d) CO32- 4 e) C2H6 -3 f) CH3OH -2 3. For each of the following reactions, identify: the oxidizing agent, the oxidized substance and the reduced substance. a) Cu2+ (aq) + Zn(s) --------> Cu(s) + Zn2+ (aq) Oxidized substance Zn Reduced substance Cu2+ Oxidizing agent Cu2+ Reducing agent Zn b) Cl2 (g) + 2 Na(s) --------> 2 Na+ (aq) + 2 Cl- (aq) Oxidized substance Na Reduced substance Cl2 Oxidizing Agent Cl2 Reducing Agent Na 4. Circle each formula that is capable of losing an O2 Cl- Fe Na + 5 electron. Determine the oxidation number of the underlined element. PbSO4 +6 ClO3- +5 HPO32- +3 Na2O2 -1 CaH2 -1 Al2(SO4)3 +6 NaIO3 +5 C4H12 -3 6. Al3+ + Zn → Al + Zn2+ +3 0 0 +2 oxidation numbers! Oxidized substance oxidizing agent Al3+ 7. Cr2O72- + ClO2- → Cr3+ + ClO4- +6 +3 +3 +7 oxidation numbers Reduced substance Cr2O72- Oxidizing agent Cr2O72- 8. Enter the oxidation number of each of the underlined elements. a) NH3 -3 b) H2SO4 +6 c) ZnCO3 +4 d) Al(OH)3 +3 e) Na 0 f) Cl2 0 9. Balance the redox equation using the mean reaction method. Al & AgNO3 Al → Al3+ + 3e- 3(Ag+ + 1e- → Ag) Al + 3Ag+ → Al3+ + 3Ag 10. Circle each formula that is capable of losing an O2 Cl- Fe Na + 11 electron. Cr3+ + ClO4- → Cr2O72- + ClO2- +3 +7 +6 +3 oxidation numbers Reduced substance ClO4- Oxidizing agent ClO4- 12. O3 + H2O + SO2 → SO42- + O2 + 2H+ ? +4 +6 0 oxidation numbers Oxidized substance SO2 REDUCing agent SO2 13. 3As2O3 + 4NO3- + 7H2O + 4 H+ → 6H3AsO4 + 4NO +3 +5 +2 Oxidation numbers Reduced substance NO3- Reducer agent As2O3 WS - 3 Spontaneous and non-spontaneous Redox reactions Describe each reaction as spontaneous or non-spontaneous. 1. Au+3 + Fe+3 ----- > Fe+2 + Au nonspontaneous (two oxidizing agents) 2. Pb + Fe+3 ------ > Fe+2 + Pb+2 spontaneously 3. Cl2 + F- ------> + 2Cl- name 4. S2O8-2 + Pb ------> 2SO4-2 + Spontaneous Pb+2 5.Cu+2 + 2Br- ------> Cu + Br2 non-spontaneous 6. Sn+2 + Br2 ------> Sn+4 + 2Br- spontaneous 7. Pb+2 + Fe+2 ------ > Fe+3 + Nonpontaneous Pb 8. 1 M HCl can be kept in an iron container. If the answer is no, write a balanced equation for the reaction that would occur. No Faith + 2H+ -------- > Fe2+ + H2 9. Can you keep 1 M HCl in an Ag container. If the answer is no, write a balanced equation for the reaction that would occur. Yes. There's no reaction. 10. Can you keep 1 M HNO3 in an Ag container. If the answer is no, write a balanced equation for the reaction that would occur. (remember that HNO3 consists of two H+ and NO3- ions) No 3Ag + NO3- + 4H+ -------- > 3Ag+ + NO + 2H2O 11. You can keep 1 M HNO3 in an Au container. If the answer is no, write a balanced equation for the reaction that would occur. (Remember, HNO3 consists of two H+ and NO3- ions) Yes. There's no reaction. 12. Circulate each formula that is capable of losing an O2 Cl- Fe Na+ 13 elecron. Determine the oxidation number of the underlined element. PbSO4 6 ClO3- 5 HP032- 3 Na2O2 -1 CaH2 -1 Al2(SO4)3 6 NaIO3 5 C4H12 -3 14. Al3+ + Zn ---------> Al Zn2+ Oxidized substance Zn Oxidizing Agent Al3+ 15. Cr2O72- + ClO2- ----------> Cr3+ + ClO4- Reduced substance Cr2O72- Oxidizing agent Cr2O72- 16. Enter the oxidation number of each of the underlined elements. a) NH3 -3 b) H2SO4 6 c) ZnCO3 4 d) Al(OH)3 3 e) Na 0 f) Cl2 0 17. Balance the redox equation using the mean reaction method. Al + 3Ag+ ----------> Al3+ + 3Ag 18. Circle of each formula that is able to lose an O2 Cl- Fe Na+ electron Determine the oxidation number for the underlined element. 19. PbsO4 2 20. ClO3- 5 21. HPO32- 3 22. Na202 -1 23. CaH2 -1 24. NaIO3 5 25. C4H12 -3 26. Al2(SO4)3 6 27. Al3+ + Zn ---------- > Al + Zn2+ Oxidized Substance Zn Oxidizing Agent Al3+ 28. Cr2O72- + ClO2- ----------> Cr3+ + ClO4- Reduced substance Cr2O72- Oxidizing agent Cr2O72- 29. O3 + H2O + SO2 ----- > SO42- + O2 + 2H+ Oxidized substance SO2 REDUCing agent SO2 30. 3As2O3 + 4NO3- + 7H2O + 4 H+ --------> 6H3AsO4 + 4NO Reduced substance NO3- Reducer agent As2O3 WS - 4 Balance Redox Reactions Balance each of the following half cell reactions. (In each case let's assume that the reaction takes place in an ACIDIC solution.) Also, indicate whether the reaction is oxidation or reduction. 1. 5H2O + S2O32- ----------> 2SO42- + 10H+ + 8e- oxidation 2. 8H+ + 5e- + MnO4- ----------> Mn2+ + 4H2O reduction 3. 4H2O + Depending ---------- > AsO43- + 8H+ + 5e- oxidation 4. 7H2O + 2Cr3+ ---------- > Cr2O72- + 14H+ + 6e- oxidation 5. 2H2O + Pb2+ ---------- > PbO2 + 4H+ + 2e- oxidation 6. 8H+ + SO42- + 6e- ----------> S + 4H2O reduction 7. 4H+ + NO3- + 5e- ---------- > N + 3H2O reduction 8. 10H+ + 8e- + NO3- ---------- > NH4+ + 3H2O reduction 9. 12H+ + 10e- + 2BrO3- Br2 + 6H2O reduction Balance Half cell Desenlafe in balance of half cell reactions in basic solution. 10. 3e- + 2H2O + NO3- ---------- > NO + 4OH- 11. 4H2O + 5e- + MnO4- ----------> Mn2+ + 8OH- 12. 2H2O + 2IVO3 + 8e- ---------- > I2 + 2VO22- + 4OH- 13. 2IVO3 + 2e- ---------- > I2 + 2VO3- 14. Pb2+ + 4OH- ----------> PbO2 + 2H2O + 2e- 15. 4H2O + 6e- + SO42- ---------- > S + 8OH- 16. 10 OH- + S2O32- ----------> 2SO42- + 5H2O + 8e- 17. 7H2O + 8e- + NO3- ---------- > NH4+ + 10 OH- 18. 6H2O + 10e- + 2BrO3- ----------> Br2 + 12 OH- 19. Determine whether each of the following changes is oxidation, reduction, or none. SO32- --------> SO42- CaO oxidation --------> Ca reduction Ca CrO42- --------> Cr2O72- ni CrO42- --------> Reduction of Cr3+ 2I- --------> oxidation I2 IO3- --------> I2 reduction MnO4- --------> Reduction of ClO2- --------> ClO- reduction 20. Cr2O72- + Fe2+ --------> Cr3+ + Fe3+ Oxidized substance Fe2+ Reduced substance Cr2O72- Oxidizing agent Cr2O72- Reducing agent Fe2+ WS #5 Balance of Redox Reactions in Acid and Basic Solution Balance each redox equation. Suppose they're all spontaneous. Use the half reaction method. 1. 2O2- + 2F2 ---------- > O2 + 4F- 2. 4Al + 3O2 ---------- > 6O2- + 4Al3+ 3. 2K + Zn+2 ---------- > Zn + 2K+ Balance each half reaction in basic solution. 4. Cr2O72- + 7H2O + 6e- ---------- > 14OH- + 2Cr3+ 5. NO + 4OH- ----------> 2H2O + NO3- + 3e- 6. 2H2O + 2e- + SO42- ----------> SO2 + 4OH- 7. 2MnO2 + H2O + 2e- ----------> Mn2O3 + 2OH- Balance each redox reaction in acid solution using the mean reaction method. 8. 8H+ + 3H2O2 + Cr2O72- -------> 3O2 + 2Cr3+ + 7H2O 9. TeO32 - + 2N2O4 + H2O ------- > Te + 4NO3- + 2H+ 10. 12H2O + 10V ------- > 6H+ + 3H2V2O42- + 4VH3 11. 8H+ + 5PbO2 + I2 ------- > 5Pb2+ + 2IO3- + 4H2O 12. 12H2O + 8A -------> 3H2AsO4- + 5AsH3 + 3H+ Balance each redox reaction into basic solution using the mean reaction method. 13. 3O2 + 8OH- + 2Cr3+ -------> H2O + 3H2O2 + Cr2O72- 14. H2O + Te + 4NO3- -------> TeO32- + 2OH- + 2N2O4 15. 10IVO3 + 28OH- -------> 10VO22- + I2 + 8IO3- 14H2O 16. 8OH- + 5Pb2+ + 2IO3- -------> 5PbO2 + I2 + 4H2O 17. 2H2O + O2 + 4Fe(H2O)62+ ----> 4Fe(H2O)63+ + 4OH- Change status represents oxidation, reduction or none (use of oxidation #s). 18. Reduction of MnO2 --------> Mn2O3 19. NH3 -------- > oxidation OF NO2 20. HClO4 ------- > HCl + H2O 21 reduction. O2 --------> O2- reduction 22. P2O5 --------- > P4H10 reduction Determine the number of oxidation 23. H2SO4 22. HSO4- 6 24. P4 0 23. NaH -1 25. UO3 6 24. Na2O2 -1 26. U2O5 5 25. PbSO4 2 WS #6 Review 1. Describe each one in his own words 1. Oxidation - loss of electrons 2. Reduction - electron gain 3. Oxidizing agent - causes oxidation when subjected to a reduction 4. Reducing agent - causes reduction when subjected to oxidation 2. Write half reactions for each. Describe as oxidation or reduction. Circle of all oxidizing agents. a) Na ----------> Na+ + e- oxidation b) Ca ----------> Ca2+ + 2e- oxidation c) Al3+ + 3e- ----------> Al reduction d) 2F1- ----------> F2 + 2e- oxidation e N2 + 6e- ----------> 2N3- reduction f) 2O2- ----------> O2 + 4e- oxidation 3. Type the reaction between the following: Use the mean reaction method. a) Ca + Al(NO3)3 3Ca + 2Al3+ ----------> 2Al + 3Ca2+ b) Sn + AgNO3 Sn + 2Ag+ ----------> 2Ag + Sn2+ c) Sn + Au(NO3)3 3Sn + 2Au3+ ----------> 2Au + 3Sn2+ 4. Circle of each reducing agent: Cu Cu+ Al Al3+ 5. Circle of each oxidizing agent: F- F O2- O2 6. Ni+2 reacts with Mn, however, Al+3 does not react with Mn. Classifies oxidizing agents in decreasing order of force. Classifies reducing agents in decreasing order of force. Stronger oxidizing agent Ni2+ + 2e- ----------> Ni Mn2+ + 2e- ----------> Mn Al3+ + 3e- ----------> The strongest reducing agent 7. Ag+ reacts with Pb, however, Ca+2 does not react with Pb. Rank reducing agents in order to decrease strength. Classifies oxidizing agents in decreasing order of force. Stronger oxidizing agent Ag+ + 1e- ----------> Ag Pb2+ + 2e- ----------> Pb Ca2+ + 2e- ----------> Strongest ca reducing agent stronger ca 8. Cl2 reacts with Ag, however, Ag does not react with Mg+2. Classifies oxidizing agents in decreasing order of force. Classifies reducing agents in decreasing order of force. Stronger oxidizing agent Cl2 + 2e- --------> 2Cl- Ag+ + 1e- ----------> Ag Mg2+ + 2e- ----------> Mg strongest reducing agent 9. Ni+2 reacts with Mn, however, Al+3 does not react with Mn. Classifies reducing agents in decreasing order of force. Classifies oxidizing agents in decreasing order of force. Stronger oxidizing agent Ni2+ + 2e- ----------> Ni Mn2+ + 2e- ----------> Mn Al3+ + ---------- > Reducing agent to the strongest Al 10. Cl2 reacts with Br-, however, I2 does not react with Br-. Classify oxidizing agents in order to decrease strength. Classifies reducing agents in decreasing order of force. Stronger oxidizing agent Cl2 + 2e- --------> 2Cl- Br2 + 2e- --------> 2Br- I2 + 2e- --------> 2I- stronger reducing agent Classify as oxidation, reduction or none. 11. Reduction SO42- --------> S2- 12. MnO2 --------> MnO4- oxidation 13. Cr2O72- --------> CrO42- or 14. Io3 Reduction- --------> I2 15. Given the following lab data SnCl2 & Ni Spontaneous Ni(NO3)2 & Fe Spontaneous Cr(NO3)3 & Fe Non spontaneous. i) Write three balanced equations. Ni + Sn2+ ----------> Ni2+ + Sn Fe + Ni2+ ----------> Fe2+ + Ni Fe + Cr3+ <---------- Fe2+ Cr ii) Classify oxidizing agents in decreasing order of force. Stronger oxidizing agent Sn2+ + 2e- ----------> Sn Ni2+ + 2e- ----------> Ni Fe2+ + 2e- ----------> Fe Cr3+ + 3e- ----------> Cr strongest reducing agent Classifies reducing agents in decreasing order of force. See above. iv) Will SnCl2 react with Cr? Explain? Yes, because Sn2+ is a stronger oxidizing agent than Cr3+. v) Will Fe2+ react with Sn? No, because Fe2+ is a weaker oxidizing agent than Sn2+ 16. 2H+ + 2MnO4- + 5H2S --------> 5S + 6H2O + 2MnO oxidizing agent reducing agent 17. 2H+ + 10SO42- + 4Br2 ----------> 5S2O32- + 8BrO3- + oxidizing agent H2O reducing agent 18. Balance the redox equation in the acidic solution IPO4 → I2 + IO3- + PO43- Track: You need two half reaction IPO4 → I2 + PO43- IPO4 → IO3- + PO43- 9H2O + 5IPO4 → 3IO3- + I2 + 5PO43- + 18H+ 19. Describe as spontaneous or non-spontaneous. Use your reduction potential chart. a) ZnCl2 & Cu nonspontaneous b) CuCl2 & NaCl nonspontaneous c) Br2 & Fe2+ spontaneous d) H2S & Al3+ nonspontaneous 20. Can you keep HCl in a Zn container? No, spontaneous reaction. What about an Au container? Yes, spontaneous reaction. Balance in basic solution 21. H2O + 10SO42- + 4Br2 ------> 5S2O32- + 2OH- + 8BrO3- Classify as oxidizing agent, reducing agent or both depending on your position in the table. Indicate the Eoor voltage of your position. Some of these are both, so indicate two voltages and indicate that it can be an oxidizing and reducing agent. E.g. MnO4- (in acid) oxidizing agent 1.51 v 22. Br2 oxidizing agent 1.09 v 23. Fe2+ oxidizing agent / reducing agent -0.45 v / - 0.77 v 24. MnO4- (water) oxidizing agent 0.60 v 25. No gearbox agent 0.26 v 26. Cr3+ Oxidizing Agent -0.74 v / -0.41v 27. Oxidizing agent H2O / reducing agent -0.41 v / -0.82 v Indicate as spontaneous or non-spontaneous. 28. Non-spontaneous MnO4- & Fe2+ 29. Cu2+ & Br- non-spontaneous 30. HNO3 & Spontaneous Ag 31. MnO4- (acid) and spontaneous H2O 32. Neither(s) and Al3+ non-spontaneous 33. Spontaneous HCl & Mg Write each mean oxidation and reduction reaction for each previous question. Determine the Eo for each. Calculate the Eo for the overall reaction. 34. MnO4- + 2H2O + 3e- --------> MnO2 + 4OH- +0.60 v 3(Fe2+ ----------> Fe3+ + 1e-) -0.77 v MnO4- + 2H2O + 3Fe2+ ----------> 3Fe3+ + MnO2 + + + + -0.17 v 35. 36. NO3- + 4H+ +3e- ----------> NO + 2H2O +0.96 v 3(Ag ----------> Ag+ + 1e-) -0.80 v NO3- + 4H+ + 3Ag ----------> NO + 2H2O + 3Ag+ +0.16 v 37. 38. 39. 2H+ + 2e- ------> H2 0.00 v Mg ----------> Mg2+ + 2e- 2.37 v Mg + 2H+ ----------> Mg2+ + H2 2.37 v WS a 7 Electrochemical Cells 1. Oxidation is

when electrons are lost. 2. Reduction is when electrons are gained. 3. The reducing agent suffers from oxidation. 4. The oxidizing agent suffers a reduction. 5. A negative voltage means that the reaction does not spontaneous. 6. In an electrochemical cell the electrons come out of the electrode, which is negative. 7. In an electrochemical cell the reduction reaction is higher in the table, while the oxidation reaction is lower. . 8. Cathode is the place of reduction and anod is the oxidation site. . 9. Anions migrate to the aodo and cations migrate to the cathode. 10. Anions have a negative charge and cations have a positive charge. Draw and fully analyze each electrochemical cell. 11. Zn / Zn(NO3)2 ‖ Cu / Cu(NO3)2 12. Ag / AgNO3 ‖ H2 / HCl WS 8 1. In an electrolyte cell, the reduction occurs in the negative electrode and oxidation occurs in the positive electrode. 2. If there are two possible reduction reactions, the highest one occurs on the chart. 3. For reduction, the chart is read from left to right. 4. For oxidation, the table is read from right to left and the voltage sign is changed. 5. If there are two possible oxidation reactions, the lowest is produced on the chart. 6. Corrosion of a metal is oxidation. 7. Electrolysis uses electrical energy. 8. Electrochemical cells produce electrical energy. 9. Electrolyte cells use electrical energy. 10. What is the standard reference cell? Eo or v Draw and fully analyze each electrolyte cell. 11. Molten NaCl Cathode: Na+ + 1e- → Na(s) -2.71 v apod: 2Cl- → Cl2 + 2e- -1.36 v Generally: 2Na+ + 2Cl- → Cl2 + 2Na(s) -4.07 v - MTV +4.07 v 12. Aqueous Na2SO4 Candido: 2H2O + 2e- → H2 + 2OH- -0.41 v apododo: H2O → 2H+ + 1/o1/2O2 + 2e- -0.82 v In general: H2O → H2 + 1/2O2 -1.23 v MTV at +1.23 v 13. K2O liquid cathode: K+ + 1e- → K(s) -2.93 v apod: 2O2- → O2 + 4e- ? v Overall: 4K+ + 2O2- → O2 + 4K(s) -? v MTV +? v 14. 1.0 M LiI Cathode: Cathode: 2H2O + 2e- → H2 + 2OH- -0.41 v apod: 2I- → I2 +2 e- -0.54 v In general: 2H2O + 2I- → I2 + H2 + 2OH- -0.95 v MTV + +0.95 v 15. 250ml of 0.200M MnO4- reacts with excess SO3-2. How many grams of MnO2 are produced? This is Chemistry 11 stoichiometry. 2MnO4- + 3SO3-2 + H2O -----> 2MnO2 + 3SO4-2 + 2OH- 0.250L MnO4- x 0.200 mol x 2 mol MnO2 x 86.9g to 4.34g L 2 mol MnO4- mol 16. Determine the oxidation number for each underlined atom. MnO2 4 Cr2O7-2 6 IO3- 5 C2O4-2 3 Al(NO3)3 5 17. Describe each term: Salt bridge- a U tube filled with salt solution that allows ions to flow into an electrochemical cell. Electrolyte- a solution that conducts anodo electricity - an electrode that is the site of Cathode oxidation - an electrode that is the site of spontaneous reduction- a reaction that occurs naturally and has a positive voltage Electron affinity - the ability of a metal to attract electrons 18. What if you used an aluminum spoon to remove a FeSO4(aq) solution? Write a reaction and calculate Eo. 2A + 3Fe2+ -------> 2Al3+ + E0 to 1.21 v Spontaneous. There would be a reaction! 19. Draw an electrochemical cell using Electrodes Cu and Ag. Cathode (+) Apod (-) Ag Cu Ag+ + 1e---------> Ag 0.80v Cu -------> Cu2 + 2e -0.34v 2Ag+ Cu ------> 2Ag + Cu2+ E0 to 0.46 v spontaneous 20. 250 ml of .500M MnO4- is required to valuate a sample of 100 ml of SO3-2. Calculate the [SO3-2] 2MnO4- + 3SO3-2 + H2O -----> 2MnO2 + 3SO4-2 + 2OH- .250L MnO4- x 0.500 mol x 3 mol SO3-2 L 2MnO4- a 1.88M 0.100L 21. How is the breathalyzer reaction used to determine the blood alcohol content (you may need to look for this in your textbook)? The breathalyzer reaction uses a spontaneous redox reaction between Cr2O72- acid and C2H5OH ethanol. If alcohol is present in the breath sample, it will react with a Cr2O72 solution- reducing the orange color as it reacts to the Cr3+ form, which is green. The drunker you are, the greater the reduction in orange, which is measured with a spectrophotometer. 22. 2H+ + Mg-----> Mg+2 +H2 Oxidizing Agent H+ Mg WS reducer agent #9 Electrolytic, Electrochemical Cells & Application Determine the mean reactions for each cell and the cell voltage or minimum theoretical voltage and general equation. 1. Electrochemical cell Ag / Pb. Anode: Pb Cathode: Ag Anode reaction: Pb --------> Pb2+ + 2e- Cathode reaction: Ag+ + 1e- -------> General reaction: Pb + 2Ag+ -----> Pb2+ + 2Ag Voltage: 0.93v 2. Electrolytic cell ZnCl2(l) Adodo: C Cathode: C Apod reaction: 2Cl- --------> Cl2 + 2e- Cathode reaction: Zn2+ + 2e- -------> Zn General reaction: 2Cl- + Zn2+ -----> Cl2 + Zn MTV: +2.12 v 3. Electrolytic cell CuSO4(aq) (electro-winning) aodo: C Cathode: C Acaode reaction: H2O --------> 2H+ + 1/2O2 + 2e- Reaction cathode: Cu2+ + 2e- -------> Cu General reaction: H2O + Cu2+ -----> 2H+ + 1/2O2 + Cu MTV: +0.48 v 4. The electrolysis of 1M NaI (electro-winner) anode: C Cathode: C Anode reaction: 2I- --------> I2 + 2e- Cathode reaction: 2H2O + 2e- -------> H2 + 2OH- Global reaction: 2H2O + 2I- -----> H2 + 2OH- + I2 MTV: +0.95 v5. The reaction needed to make Al. The electrolyte is Al2O3 and its phase is molten (melted or aqueous). To lower the mp. cryolite is used from 2000 oC to 800 oC. Adodo: C Cathode: C Apod reaction: 2O2- -------> O2 + 4e- Cathode reaction: Al3+ + 3e- -------> To General reaction: + 4Al3+ -----> 3O2 + 4A 6. The reaction needed to electroplate a penny of copper with silver. Anode: Ag Cathode: penny Anode reaction: Ag-----> Ag+ + e- Cathode reaction: Ag+ + e- -----> Ag 7. The necessary reaction to the nickel plate a penny of copper. Alodo: Neither Cathode: penny Adoode reaction: Ni-----> Ni+2 + 2e- Cathode reaction: Ni2 + 2e- -----> Ni Possible electrolyte Ni(NO3)2 8. The reaction used in the electrore refine the lead. Anode: Impure Lead Cathode: Pure Lead Anode reaction: Pb-----> Pb+2 + 2e- Cathode reaction: Pb2+ + 2e- -----> Pb WS - 10 Electrolytic, Electrochemical Cells, Corrosion, & Cathodic Protection Determine the mean reactions for each cell and the minimum theoretical cell or voltage voltage. 1. Zn / Mg electrochemical ado cell: Mg cathode: Zn aanodo reaction: Mg --------> Mg2+ + 2e- Cathode reaction: Zn + 2e- -------> Zn General reaction: Mg + Zn2+ -----> Mg2 + Zn Voltage: 1.61v 2. The electrolyte cell used to produce Al. Electrolyte: Phase Al2O3 (aqueous or molten) Adodo: C Cathode: C Apod reaction: 2O2- --------> O2 + 4e- Cathode reaction: Al3+ + 3e- -------> To General reaction: 6O2- + 4Al3+ -----> 3O2 + 4Al 3. Electrolysis KI(aq) adodo: Cathode C: C Aodo reaction: 2I- --------> I2 + 2e- Cathode reaction: 2H2O + 2e- -------> H2 + 2OH- General reaction: 2H2O + 2I- -----> H2 + 2OH- + I2 MTV: +0.95 v 4. The electroresine of Pb anodo: Impure lead cathode: Pure lead anod reaction: Pb-----> Pb+2 + 2e- Cathode reaction: Pb2+ + 2e- -----> Pb 5. Nickel plated an iron nail. Adodo: Neither Cathode: nail ado reaction: Ni-----> Ni+2 + 2e- Cathode reaction: Ni2+ + 2e- -----> Ni Possible electrolyte Ni(NO3)2 The -ve side of the power supply is connected to the nail 6. Draw an electrochemical cell Ag/Zn. Anodo: Zn Cathode: Ag Anode Zn --------> Zn2+ + 2e- Cathode reaction: Ag+ + 1e- -------> Ag General reaction: Zn + 2Ag+ -----> Zn2+ + 2Ag Voltage: 1.56v 7. Draw an electrolytic cell KF(l). Adodo: C Cathode: C Acaodo reaction: 2F- --------> F2 + 2e- Cathode reaction: K+ + e- --------> K General reaction: 2F- + 2K+ -----> Cl2 + K MTV: +5.80v 8. Draw an electrolytic cell KF(aq). Adodo: C Cathode: C Adoode reaction: H2O --------> 2H+ + 1/2O2 + 2e- Cathode reaction: 2H2O + 2e- -------> H2 + 2OH- General reaction: H2O -----> H2 + 1/2O2 MTV: +1.23 v 9. Draw a FeI2(aq) electrolytic cell. Adodo: C Cathode: C Adoodo reaction: 2I- --------> I2 + 2e- Cathode reaction: Fe2+ + 2e- -------> Faith General reaction: Fe2+ + 2I- -----> Fe + I2 MTV: +0.99 v 10. Draw a Cd/Pb electrochemical cell. Cd is not on the reduction chart, however, the Cd electrode gains mass and total cellular potential is .5v. Determine half cell potential for Cd. aodo: Pb Cathode: Cd aode reaction: Pb --------> Pb2+ + 2e- 0.13v Cathode reaction: Cd+2 + 2e- -------> Zn x volts General reaction: Pb + Cd2+ -----> Pb2+ + Cd Voltage: 0.50v 0.13 + x at 0.50 x at 0.37v 11. 2HIO3 + 5H2SO3 ----------> I2 + 5H2SO4 + Oxidizing Agent H2O Oxidized Substance H2SO3 Reduced REDUCing AGENT HIO3 H2SO3 12. What is the fuel in a fuel cell? H2 and O2 13. Describe the differences and similarities between an electrolyte and electrochemical cell. Electrolytic Electrochemical Uses Electricity Produces Spontaneous Electricity Nopontanea Causes Chemicals Uses Chemicals Inert Carbon Electrodes It usually has a salt bridge The negative electrode is the reduction The highest metal is reduction Oxidation occurs in the anodo and the reduction occurs in the cathode. Anions migrate to the aodo and cations migrate to the cathode. Electrons range from acaodo to cathode through wire. 14. Describe and give two examples of electroganing. Water electrolysis to make H2 and O2. The electrolysis of Al2O3 to make Al and O2. 15. Describe and give an example of electrorefinning. The electrorefinning of Pb. 16. List three metals that can be gained from the aqueous solution. Pb Au Ag Zn Cu Sn 17. List three metals that cannot be gained with an aqueous solution. Na K Li Ca Mg Al 18. Indicate two metals that can be used to cathodically protect Fe. Describe how they protect iron from corrosion. Zn and Mg. When they join Fe they form an electrochemical cell. Zn or Mg is a stronger reducing agent (lower on the chart) and is the ado and Fe is the cathode. Since cathode is the place of reduction, Faith cannot oxidize or corrode. 19. Write the mean reaction that describes iron corrosion. Fe -------- > Fe2+ +2e- 20. Write down the mean reaction that describes the reduction reaction that occurs when iron corrodes in the air and water. 2e- + H2O + 1/2O2 ---------- > 2OH- 21. Why does iron corrode faster in salt water? Salt acts as a salt bridge and increases the reaction rate in an electrochemical cell. 22. Write the aodo and cathode reaction in an electrolyte cell with a CaCl2 electrolyte (l). Cathode: Ca2+ + 2e- ---------> Ca aodo: 2Cl- ----------> Cl2 + 2e- 23. Explain why would Zn or Cu choose to cathodically protect iron? Zn. It is a stronger reducing agent than Faith and will allow Faith to be the cathode, which cannot corrode. 24. Choose a redox reagent suitable for oxidizing Cl- a ClO4- in a redox rating. MnO4- in acid gives a spontaneous reaction, as well as a color change from purple to light. 25. Describe as an electrochemical or electrolyte cell: a) Electrochemical fuel cell b)Charging an electrolytic car battery c) Discharge of an electrochemical car battery d) Or electrolytic plated e) Industrial To electrolytic production f) Cl2 electrolytic production 26) Al and AgNO3(aq) are mixed and the surface of the Al darkens. List the two oxidizing agents in the decreasing force. List the two reducing agents in decreasing force. Oxidizing Agents Ag+ Al3+ Reducing Agents Al Ag 29. Discuss this label for each aodo and cathode. See Diagram Write each aodo and cathode reaction. See Diagram Indicate ion migration in each cell. See Diagram Determine the initial voltage of the electrochemical cell. See Diagram Determine MTV for electrolyte cell. See Diagram Will electrolysis occur? Yes Indicates the flow of electrons. See Diagram Indicate all electrodes that gain mass. Ag Cu Indicate all electrodes that lose mass. Mg What about [NO3-] in the Mean Mg Cell? Increases What happens to [Ag+] in the Mean Ag Cell? Decreases What about [Mg2+] in the mean Mg cell? Increases What is the potential for equilibrium electrochemical cells? 0 V What chemical is made in the Pt electrode on the right? Cu What chemicals are manufactured in the Pt electrode on the left? O2 and H+