Balancing redox equations worksheet answers chemistry if8766

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NOTE: This worksheet is intended for Honors/Pre-AP students. **Teacher Notes**

Name **Key** Class ______ Date ____________

Balancing Redox Equations

Rules for assigning oxidation numbers (see Holt p. 591):

1. **The oxidation number of any uncombined element is 0.**
2. **The oxidation number of a monatomic ion equals the charge on the ion.**
3. **The more electronegative element in a binary compound is assigned the number equal to the charge it would have if it were an ion.**
4. **The oxidation number of fluorine is a compound is always -1.**
5. **Oxygen has an oxidation number of -2 unless it is combined with F, when it is +2, or it is in a peroxide, such as H_2O_2, when it is -1.**
6. **The oxidation state of hydrogen in most compounds is +1 unless it is combined with a metal, in which case it is -1.**
7. **In compounds, Group 1 and 2 elements and aluminum have oxidation numbers of +1, +2, and +3, respectively.**
8. **The sum of the oxidation numbers of all atoms in a neutral compound is 0.**
9. **The sum of the oxidation numbers of all atoms in a polyatomic ion equals the charge of the ion.**

1. Determine the oxidation state of each element in the following:

   a. P_2O_5 **P: +5, O: -2**
   b. ZnS **Zn: +2, S: -2**
   c. $CuSO_4$ **Cu: +2, S: +6, O: -2**
   d. PO_4^{-3} **P: +5, O: -2**
   e. $(NH_4)_2SO_4$ **N: -3, H: +1, S: +6, O: -2**
   f. Ag_3PO_3 **Ag: +1, P: +3, O: -2**
   g. $HgCl_2$ **Hg: +2, Cl: -1**
   h. $MnCO_3$ **Mn: +2, C: +4, O: -2**
   i. I_2O_7 **I: +7, O: -2**
   j. $Fe(MnO_4)_3$ **Fe: +3, Mn: +7, O: -2**

Definitions:

- **A species whose oxidation number increases is oxidized.**
- **A species whose oxidation number decreases is reduced.**
- **The reduced substance is called the oxidizing agent.**
- **The oxidized substance is called the reducing agent.**

2. Determine the element being oxidized and reduced in each of the following equations:

   0 +1+4-2 +1-2 +1-1 +1+6-2
   a. $I_2 + H_2SO_3 + H_2O \rightarrow HI + H_2SO_4$
   Element being oxidized: **sulfur** Element being reduced: **iodine**

8) 1 C_3H_8 + 5 $O_2 \rightarrow$ 3 CO_2 + 4 H_2O
9) 2 C_8H_{18} + 25 $O_2 \rightarrow$ 16 CO_2 + 18 H_2O
10) 1 $FeCl_3$ + 3 NaOH $\rightarrow$ 1 $Fe(OH)_3$ + 3 NaCl
11) 4 P + 5 $O_2 \rightarrow$ 2 P_2O_5
12) 2 Na + 2 $H_2O \rightarrow$ 2 NaOH + 1 H_2
13) 2 $Ag_2O \rightarrow$ 4 Ag + 1 O_2
14) 1 S_8 + 12 $O_2 \rightarrow$ 8 SO_3
15) 6 CO_2 + 6 $H_2O \rightarrow$ 1 $C_6H_{12}O_6$ + 6 O_2
16) 2 K + 1 $MgBr_2 \rightarrow$ 2 KBr + 1 Mg

Mg^{2+}, Br^- ERROR $MgBr_2$

17) 2 HCl + 1 $CaCO_3 \rightarrow$ 1 $CaCl_2$ + 1 H_2O + 1 CO_2
18) 1 HNO_3 + 1 $NaHCO_3 \rightarrow$ 1 $NaNO_3$ + 1 H_2O + 1 CO_2
19) 2 H_2O + 1 $O_2 \rightarrow$ 2 H_2O_2
20) 2 NaBr + 1 $CaF_2 \rightarrow$ 2 NaF + 1 $CaBr_2$

O I L — reducing agent   R I G — oxidizing agent

REDOX REACTIONS

Name ____________

For the equations below, identify the substance oxidized, the substance reduced, the oxidizing agent, the reducing agent, and write the oxidation and reduction half reactions.

Example: oxidized Mg + reduced $Br_2 \rightarrow MgBr_2$ (Mg: reducing agent; Br_2: oxidizing agent)
oxidation half reaction: $Mg^0 \rightarrow Mg^{+2} + 2e^-$
reduction half reaction: $2e^- + Br_2^0 \rightarrow 2Br^-$

1.	$2H_2$ (ra) + O_2 (oa) $\rightarrow 2H_2O$	oxid $2H_2 \rightarrow 4H^+ + 4e^-$ red $O_2 + 4e^- \rightarrow 2O^{2-}$
2.	Fe (ra) + Zn^{2+} (oa) $\rightarrow Fe^{2+}$ + Zn	oxid $Fe \rightarrow Fe^{2+} + 2e^-$ red $Zn^{2+} + 2e^- \rightarrow Zn$
3.	2Al (ra) + $3Fe^{+2}$ (oa) $\rightarrow 2Al^{+3}$ + 3Fe	oxid $2Al \rightarrow 2Al^{3+} + 6e^-$ red $3Fe^{2+} + 6e^- \rightarrow 3Fe$
4.	Cu (ra) + $2AgNO_3$ (oa) $\rightarrow Cu(NO_3)_2$ + 2Ag	oxid $Cu \rightarrow Cu^{2+} + 2e^-$ red $2e^- + 2Ag^+ \rightarrow 2Ag$

Name: ______________________ **Date:** __________

Balancing Chemical Equations

Balance the following chemical equations.

1. 2 Fe + 3 $H_2SO_4 \rightarrow$ 1 $Fe_2(SO_4)_3$ + 3 H_2
2. 1 CH_4 + 2 $O_2 \rightarrow$ 1 CO_2 + 2 H_2O
3. 1 $SiCl_4(\ell)$ + 2 $H_2O(\ell) \rightarrow$ 1 $SiO_2(s)$ + 4 HCl(aq)
4. 2 AgI + 1 $Na_2S \rightarrow$ 1 Ag_2S + 2 NaI
5. 4 NH_3 + 5 $O_2 \rightarrow$ 4 NO + 6 H_2O
6. 1 $FeO_3(s)$ + 3 CO(g) $\rightarrow$ 1 $Fe(\ell)$ + 3 $CO_2(g)$
7. 1 SiO_2 + 4 HF $\rightarrow$ 1 SiF_4 + 2 H_2O
8. 2 NaBr + 1 $Cl_2 \rightarrow$ 2 NaCl + 1 Br_2
9. 4 $(NH_4)_3PO_4$ + 3 $Pb(NO_3)_4 \rightarrow$ 1 $Pb_3(PO_4)_4$ + 12 NH_4NO_3
10. 1 $Mg(OH)_2$ + 2 HCl $\rightarrow$ 1 $MgCl_2$ + 2 H_2O

Balancing Chemical Equations Worksheet

1. _____ H_2 + _____ O_2 → _____ H_2O
2. _____ N_2 + _____ H_2 → _____ NH_3
3. _____ S_8 + _____ O_2 → _____ SO_3
4. _____ N_2 + _____ O_2 → _____ N_2O
5. _____ HgO → _____ Hg + _____ O_2
6. _____ CO_2 + _____ H_2O → _____ $C_6H_{12}O_6$ + _____ O_2
7. _____ Zn + _____ HCl → _____ $ZnCl_2$ + _____ H_2
8. _____ $SiCl_4$ + _____ H_2O → _____ H_4SiO_4 + _____ HCl
9. _____ Na + _____ H_2O → _____ $NaOH$ + _____ H_2
10. _____ H_3PO_4 → _____ $H_4P_2O_7$ + _____ H_2O
11. _____ $C_{10}H_{16}$ + _____ Cl_2 → _____ C + _____ HCl
12. _____ CO_2 + _____ NH_3 → _____ $OC(NH_2)_2$ + _____ H_2O
13. _____ Si_2H_3 + _____ O_2 → _____ SiO_2 + _____ H_2O
14. _____ $Al(OH)_3$ + _____ H_2SO_4 → _____ $Al_2(SO_4)_3$ + _____ H_2O
15. _____ Fe + _____ O_2 → _____ Fe_2O_3
16. _____ $Fe_2(SO_4)_3$ + _____ KOH → _____ K_2SO_4 + _____ $Fe(OH)_3$
17. _____ $C_7H_6O_2$ + _____ O_2 → _____ CO_2 + _____ H_2O
18. _____ H_2SO_4 + _____ HI → _____ H_2S + _____ I_2 + _____ H_2O
19. _____ FeS_2 + _____ O_2 → _____ Fe_2O_3 + _____ SO_2
20. _____ Al + _____ FeO → _____ Al_2O_3 + _____ Fe
21. _____ Fe_2O_3 + _____ H_2 → _____ Fe + _____ H_2O
22. _____ Na_2CO_3 + _____ HCl → _____ $NaCl$ + _____ H_2O + _____ CO_2
23. _____ K + _____ Br_2 → _____ KBr
24. _____ C_7H_{16} + _____ O_2 → _____ CO_2 + _____ H_2O
25. _____ P_4 + _____ O_2 → _____ P_2O_5

How to balance an equation chemistry. Balancing a chemical equation worksheet. How do you balance chemical equations examples. How to balance chemical formulas.

1. MOLES AND MASS Name Determine the number of moles in each of the quantities below. . 25gofNaCI 2. 125gofH 2 SO ¿ 3. TOO. gofKMnO, 4. 74gofKCI 5. 35 g of CuSCVShLO Determine the number of grams in each of the quantities below. 1. 2.5 moles of NaCl 2. 0.50 moles of H2SO4 3. 1.70 moles of KMnO, 4. 0.25 moles of KCI 5, 3,2 moles of CuSO4«5H2O 50 ©Instructional Fair, Inc 2. THE MOLE AND AVOGADRO'S NUMBER One mole of a substance contains Avogadro's Number (6.02 x 1023) of molecules. How many molecules are in the quantlties below? 1. 2.0 moles 2. 1.5 moles 3. 0.75 mole 4. 15 moles 5. 0.35 mole How many moles are in the number of molecules below? 1. 6.02 xlO 23 2. 1.204x10 24 3. l.SxlO 2 0 4. 3,4xlO 2 6 5. 7 , 5 x l O] 52 ©Instructlonal Fair,Inc. 3. MIXED MOLE PROBLEMS Ñame Solve the following problems. 1. How many grams are there ¡n 1.5 x 1025 molecules of CO2? 2. What volume would the CO2 in Probiem 1 occupy at STP? 3. A sample of NH, gas occupies 75.0 liters at STP. How many molecules ¡s this? 4, What ¡s the mass of the sample of NK in Probiem 3? 5. How many atoms are there in 1.3 x 1022 molecules of NO2? ó. A 5.0 g sample of O2 ¡s in a container at STP. What volume is the container? 7. How many molecules of O2 are in the container in Probiem 6? How many atoms of oxygen? :hemlstry IF8766 53 ©Instructional Faír, Inc. 4. BALANCING CHEMICAL EQUATIONS Name Rewrite and balance the equations below. 1. Nn + H, NH. 2. KCIO, KCI + 0 2 _ 3. NaCI + F, NaF + CL 4. H2 + 02 -» H20 5. AgNO, + MgCL AgCI + Mg(NOJ 3'2 ó. AIBr3 + KBr + AI2(SOd)3 7. CH4 + 02 -» CO2 + H2O 8. C3H8 + O2 -» CO2 + H2O. 9. C8H18 ^ 10. FeCI3 + NaOH -» Fe(OH)3 + NaCI. 11. P + O, 12. Na + H2O NaOH + H, 3. Ag,O -* Ag + O, ^ 14. S0 -H O, SO, 15. CO2 + H20 -» C6H)2O6 + O2. 16. K + MgBr, KBr + Mg 17. HCI + CaCO, CaCI2 + H2O + CO2 Chemistry IF8766 58 ©Instructional Fair. 5. WORD EQUATIONS Ñame rite the word equations below as chemical equations and balance. . zinc + lead (II) nitrate yield zinc nitrate + lead . aluminum bromide + chlorine yield aluminum chloride + bromine . sodium phosphate + calcium chloride yield calcium phosphate + sodium chloride . potassium chlorate when heated yields potassium chloride + oxygen gas . aluminum + hydrochloric acid yield aluminum chloride + hydrogen gas . calcium hydroxide + phosphoric acid yield calcium phosphate + water . copper + sulfuric acid yield copper (II) sulfate + water + sulfur dioxide . hydrogen + nitrogen monoxide yield water + nitrogen ímistry IF8766 59 ©Instructional Fair, Inc. 6. >REDICTING PRODUCTS Ñame >F CHEMICAL REACTIONS redict the products of the reactions below. Then, write the balanced equation and lassifythereaction. 1. magnesium bromide + chlorine 2. aluminum + iron (III) oxide 3. silver nitrate + zincchloride 4. hydrogen peroxide (catalyzed by manganese dioxide) 5. zinc + hydrochlorlc acid 6. sulfuric acid + sodlum hydroxide 7. sodlum + hydrogen 8. acetic acid + copper :h©mlstry IF8766 61 ©Instructional Fair, Inc. 7. CLASSIFICATION OF Ñame CHEMICAL REACTIONS Classtfy the reactions below as synthesis, decomposition, single replacement (cationic or anionic) or double replacement. . 2H2 + O2 -» 2H20 2. 2H2O -* 2H2 + O2 3. Zn + H2SOd -» ZnSOd + H2 4. 2CO + O2 -> 2CO2 5. 2HgO -» 2Hg + O, ó. 2KBr + CI2 -» 2KCI + Br2 7. CaO + H2O -» Ca(OH)2 8. AgNO, + NaCI -> AgCI + NaNCX 9. 2H2O2 -» 2H20 + 02 10. Ca(OH)2 + H2S04 -> CaSO, + 2hLO 4 "" 2 Chemistry IF8766 60 ©Instructlonal Far. 8. STOICHIOMETRY: Ñame MOLE-MOLE PROBLEMS 1. N2 ' 3H V1'2 2NH, 3 How many moles of hydrogen are needed to completely react with two moles of nitrogen? 2. 2KCIO 2KCI + 3O, How many moles of oxygen are produced by the decomposition of six moles of potassium chlorate? 3. Zn + 2HCI ZnCI'2 + H'2 ' How many moles of hydrogen are produced from the reaction of three moles of zinc with an excess of hydrochloric acid? 4, 5O, 2 3CO2 4H2O How many moles of oxygen are necessary to react completely with four moles of propane (C3H8)? 5. K3P04 AKN03)3 - 3KNCX + AIRO, How many moles of potassium nitrate are produced when two moles of potassium phosphate react with two moles of aluminum nitrate? Chemistry IF8766 62 ©Instructional Fair, md 9. STOICHIOMETRY: Ñame MASS-MASS PROBLEMS 2KCIO, 2KCI + 3O, How many grams of potassium chioride are produced if 25 g of potassium chloratt decampóse? 2. N2 + 3H2 -> 2NH3 How many grams of hydrogen are necessary to react compietely with 50.0 g of nitrogen in the above reaction? 3. How many grams of ammonia are produced in the reaction in Problem 2? 4. 2AgNO3 + BaCI2 -> 2AgCI + Ba(NO3)2 How many grams of silver chioride are produced from 5.0 g of silver nitrate reacting with an excess of barium chioride? 5. How much barium chioride is necessary to react with the silver nitrate in Problem 4? Chemlstry IF8766 64 ©Instructlonal Fair, ncj 10. STOICHIOMETRY: Ñame MIXED PROBLEMS . N2 + 3H2 -» 2NH3 What volume of NH, at STP is produced ¡f 25.0 g of N2 ¡s reacted with an excess ofH 2 ? 2. 2KCIO3 -» 2KCI + 3O2 If 5.0 g of KCIO3 is decomposed, what volume of O2 is produced at STP? 3. How many grams of KCI are produced in Probiem 2? 4. Zn + 2HCI -» ZnCI2 + H2 What volume of hydrogen at STP ¡s produced when 2.5 g of zinc react with an excess of hydrochloric acid? 5. H2SO4 2NaOH H2O + Na2SO4 How many molecules of water are produced if 2.0 g of sodium sulfate are produced in the above reaction? ó. 2AICI 2AI + 3CI, If 10.0 g of aluminum chioride are decomposed, how many molecules of CL are produced? :hemlstry IF8766 65 ©Instructíona! Fair, tnc, 11. STOICHIOMETRY: Ñame LIMITING REAGENT N2 + 3H2 -» 2NH3 How many grams of NH3 can be produced from the reaction of 28 g of N2 and 25 g of H2? 2. How much of the excess reagent ¡n Problem 1 ¡s left over? 3. Mg + 2HCI -* MgCI2 + H2 What volume of hydrogen at STP is produced from the reaction of 50.0 g of Mg and the equivalent of 75 g of HCI? 4. How much of the excess reagent in Problem 3 ¡s left over? 5. 3AgNO3 + Na3PO4 -» Ag3POd + 3NaNO3 Silver nitrate and sodium phosphate are reacted ¡n equai amounts of 200. g each. How many grams of silver phosphate are produced? 6. How much of the excess reagent ¡n Problem 5 ¡s left? Chemistry IF8766 66 ©InstructionalFairJ Balancing redox reactions FacebookTwitterLinkedInPinterest Oxidation-reduction reactions, also called redox reactions, involve the transfer of electrons from one species to another. These kinds of reactions are at the heart of energy producing devices such as batteries and fuel cells. They are also involved in many electrochemical processes by which we obtain useful materials. A reaction in which one species transfers electrons to another is called an oxidation-reduction reaction, also called a redox reaction. For example, we can think of the reaction of metallic iron with chlorine gas to form ionic iron(III) chloride as the net transfer of six electrons from two iron atoms to three chlorine molecules: Table \(\PageIndex{1}\): Redox Reaction Example Multiplier Half Reaction Direction of Electrons REDOX Reaction 2 \(Fe^0 \rightarrow Fe^{3+} + 3e^–\) electrons "pushed" oxidation 3 \(Cl_2^0 + 2e^– \rightarrow 2Cl^–\) electrons "pulled" reduction Adding the scaled reactants and scaled products results in \[2Fe^0 + 3 Cl_2^0 + \cancel{6 e^+} \rightarrow 2Fe^{3+} + \cancel{6e^-} + 2Cl^-\] and canceling electrons results in the final redox reacton is then \[2Fe(s) + 3Cl_2(g) \rightarrow 2Fe^{3+} +6Cl^- \] In essence, the Fe "pushes" electrons and the Cl2 "pulls" electrons, thereby effecting electron transfer. On this basis, we have the following definitions: Oxidation - loss of electrons by a substance Reduction - gain of electrons by a substance As this example shows, we can separate the overall redox reaction into two half reactions, one for the oxidation and one for the reduction. Notice that in the oxidation half reaction, the electrons appear on the right, and in the reduction half reaction they appear on the left. Each half reaction is multiplied by a factor so that the number of electrons produced by the oxidation is equal to the number consumed by the reduction. Oxidation and reduction always involve transfer of electrons. Therefore, there is never oxidation without reduction and vice versa in a redox reaction. Oxidizing something must cause something else to be reduced and vice versa. Therefore, the substance oxidized is seen to be the agent of the other substance's reduction, and the substance reduced is seen to be the agent of the other substance's oxidation. This leads to the following definitions: Oxidizing agent (oxidant) - a substance that causes another substance to be oxidized and is itself reduced. Reducing agent (reductant) - a substance that causes another substance to be reduced and is itself oxidized. In these terms, all redox reactions take on the general form \[Ox_1 + Red_2 \rightarrow Red_1 + Ox_2\] In this general representation, Ox1 and Ox2 are oxidizing agents (oxidants), and Red1 and Red2 are reducing agents (reductants). Thus, when Ox1 reacts with Red2, it becomes its reduced species, Red1, while at the same time Red2 becomes its oxidized species, Ox2. The process of \(Ox_1 \rightarrow Red_1\) is a reduction that might require, say, \(n\) electrons. Thus we could write this as the reduction half reaction \[Ox_1 + ne^– \rightarrow Red_1 \label{half1}\] Likewise, the process of \(Red_2 \rightarrow Ox_2\) is an oxidation that might require, say, m electrons. Thus we could write this as the oxidation half reaction \[Red_2 \rightarrow Ox_2 + me^– \label{half2}\] To write the balanced redox reaction, we want to put these two half reactions (Equations \(\ref{half1}\) and \(\ref{half2}\)) together in such a way that no net electrons show on either side of the overall reaction equation. Assuming that \(n eq m\), we will need to multiply the reduction half reaction by the factor m and the oxidation half reaction by the factor n, so that on addition the total number of electrons on both sides cancel out. Thus, we will bring our two half reactions together as follows: Table \(\PageIndex{2}\): General Redox Reaction Balancing Multiplier Half Reaction Direction of Electrons REDOX Reaction \(m\) \(Ox_1 + ne^– \rightarrow Red_1\) electrons "pushed" reduction \(n\) \(Red_2 \rightarrow Ox_2 + me^–\) electrons "pulled" oxidation Adding the scaled reactants and scaled products alogn with canceling the electrons results in to balanced redox reaction \[mOx_1 + nRed_2 \rightarrow mRed_1 + nOx_2\] Notice that by multiplying the first half reaction by \(m\) and the second by \(n\), we balanced the overall redox reaction in terms of a transfer of nm electrons, but those electrons do not show in the final balanced equation. This should always be the case for a balanced redox reaction. Redox also causes a change in the oxidation numbers of the reductant and oxidant. In a reduction, one element in a species experiences a lowering of its oxidation number, while in an oxidation the opposite occurs. This is demonstrated by expanind Table \(\PageIndex{1}\). Table \(\PageIndex{3}\): Redox Reaction Example Multiplier Half Reaction Change in Oxidation Number Reaction 2 \(Fe^0 \rightarrow Fe^{3+} + 3e^–\) Fe oxidation number increases \(0 \rightarrow +3\) Oxidation 3 \(Cl_2^0 + 2e^– \rightarrow 2Cl^–\) Cl oxidation number decreases \(0 \rightarrow -1\) Reduction The two half reactions in Table \(\PageIndex{1}\) also illustrate another important feature of balancing redox reactions. Notice that in each half reaction there is a balance both in the numbers of atoms of each kind and in the overall charge on each side. For example, in the oxidation \[Fe^0 \rightarrow Fe^{3+} + 3e^–\] we have one iron atom on each side, but also the zero charge on the left is balanced by the 3 + (3–) sum on the right from the Fe3+ ion and the three negative electrons. Likewise, in the reduction \[Cl_2^0 + 2e^– \rightarrow 2Cl^–\] the –2 charge on the left from the two electrons is balanced by the (2)(–1) charge on the right from the two chloride ions. In every half reaction and every overall redox equation there must be both a mass balance and a charge balance. For each of the following, separate the skeletal (unbalanced) equation into two half reactions. For each half reaction, balance the elements (mass balance), and then add electrons to the right or left side to make a net charge balance. Identify which half reaction is the oxidation and which is the reduction. Then, multiply each half reaction by an appropriate factor so that the two multiplied half reactions add together to make a balanced redox equation. \(Hg_2^{2+} + S_2O_3^{2-} \rightarrow Hg + S_4O_6^{2–}\) \(Al + Cr^{3+} \rightarrow Al^{3+} + Cr^{2+}\) \(Au^{3+} + I^– \rightarrow Au + I_2\) There are two principal methods for balancing redox equations: oxidation state method ion-electron method. The latter is easier to use with redox reactions in aqueous solution and if necessary can be adapted to many situations that are not in aqueous solution. Our primary interest will be in aqueous-solution redox; therefore, we will use the ion-electron method. One of the major advantages of this method is that it makes it completely unnecessary to assign individual oxidation numbers. To balance a redox equation by the ion-electron method, carry out the following steps in this sequence: Separate the skeletal equation into two half reactions. One half reaction will be a reduction and the other will be an oxidation. It is not necessary at this stage to identify which is which. Balance each half reaction separately. Balance atoms on each side of a half reaction by inspection. If the reaction occurs in acidic medium, you may add H2O and/or H+ to balance oxygen and/or hydrogen. If the reaction occurs in basic medium, you may add H2O and/or OH- to balance oxygen and/or hydrogen. Do not add any other new species (e.g., O2, H2) unless already a part of the skeletal half reaction. Balance the net charge across each half reaction by adding electrons to the side with the more positive net ionic charge. If by this process electrons are added on the left side of a half reaction, the half reaction is a reduction. If electrons are added to the right side, the half reaction is an oxidation. (If you add electrons to the same side in both half reactions, something is wrong!) Multiply both half-reactions by appropriate whole number factors, so that the number of electrons is the same in both half reactions and will cancel when the two are added together. Add the two multiplied half reactions together to obtain the overall redox equation. Check the balance. No electrons should appear in the overall redox equation. Not only should there be an element-by-element balance across the equation, but also the net charge (the sum of both ionic charges and electron charges) on both sides of the equation should be equal. Note that this procedure does not involve assigning oxidation numbers. Nonetheless, if oxidation numbers are assigned to the balanced equation, it will always occur that the reduction involves lowering an oxidation state of some element, and the oxidation involves raising an oxidation state of some element. The following examples illustrate the ion-electron procedure, starting from the skeletal equation in either acidic or basic solution. Balance \(NO_3^– + Fe^{2+} \rightarrow HNO_2 + Fe^{3+}\) in acid aqueous solution. The fact that this reaction occurs in acid aqueous solution suggest that water or \(H^+\) may be involved in the reaction. This is not a requirement from the question of course. The protonation of a species (\(NO_3^- \rightarrow HNO_2\)) further suggest this. Table \(\PageIndex{4}\): Redox Reaction Example Multiplier Half Reaction Change in Oxidation Number Reaction 2 \(Fe^{2+} \rightarrow Fe^{3+} + e^–\) Fe oxidation number increases \(+2 \rightarrow +3\) Oxidation 3 \(2e^– + 3 H^+ + NO_3^– \rightarrow HNO_2 + H_2O\) N oxidation number decreases \(+5 \rightarrow +3\) Reduction Adding these half reactions together and canceling electrons results in \[2 Fe^{2+} + 3 H^+ + NO_3^– \rightarrow 2 Fe^{3+} + HNO_2 + H_2O \] To make the oxygen balance in the \(NO_3^–/HNO_2\) reduction half reaction, we added \(H_2O\) to the right and then added \(3 H^+\) to make the hydrogen balance. These are the only allowable species to use in acid medium. Balance \(NiO_2 + Cd \rightarrow Ni(OH)_2 + Cd(OH)_2\) in basic aqueous solution. This is a "NiCad" secondary battery reaction. The fact that this reaction occurs in basic aqueous solution suggest that water or \(OH^-\) may be involved in the reaction. This is not a requirement from the question of course. Table \(\PageIndex{5}\): Redox Reaction Example Multiplier Half Reaction Change in Oxidation Number Reaction 1 \(2 OH^– + Cd \rightarrow Cd(OH)_2 + 2e^–\) Fe oxidation number increases \(+2 \rightarrow +3\) Oxidation 1 \(2e^– + 2 H_2O + NiO_2 \rightarrow Ni(OH)_2 + 2 OH^–\) Ni oxidation number decreases \(+5 \rightarrow +3\) Reduction Adding these half reactions together and canceling electrons results in \[NiO_2 + Cd + H_2O \rightarrow Ni(OH)_2 + Cd(OH)_2 \] Because this is in base, we can only add H2O and/or OH– to make the oxygen and hydrogen balances. The need to add OH– in the Cd/Cd(OH)2 half reaction is straightforward. In the NiO2/Ni half reaction, think of H2O as an acid neutralizing basic NiO2. Thus, we add two H2O to the left to neutralize the two O2– ions of NiO2, and then we add two OH– to the right side to complete the balance. Balancing oxygen and hydrogen in basic redox reactions sometimes can be difficult, because both OH– and H2O contain both elements. A trick to get around this is to balance any troublesome half-reaction or the entire redox reaction first as if it were in acid, using H+ and H2O. Then, the acid-balanced equation is converted to its form in basic medium by adding the same number of OH– to both sides of the equation that would be needed to "neutralize" any H+ in the acid-balanced equation. Combine H+ and OH– pairs to become H2O; i.e., \(H^+ + OH^– \rightleftharpoons H_2O\). The following example shows this technique for a redox reaction to be balanced in base. Balance \(I^– + MnO_4^2– \rightarrow IO_3^– + MnO_2\) in basic aqueous solution. We will balance this in acid first, then "neutralize" any \(H^+\) to convert the redox reaction to basic conditions. Table \(\PageIndex{6}\): Redox Reaction Example Multiplier Half Reaction Change in Oxidation Number Reaction 1 \(3 H_2O + I^– \rightarrow IO_3^– + 6 H^+ + 6e^–\) I oxidation number increases \(-1 \rightarrow +5\) Oxidation 1 \(2e^– + 2 H_2O + NiO_2 \rightarrow Ni(OH)_2 + 2 OH^–\) Ni oxidation number decreases \(+4 \rightarrow +2\) Reduction Adding these half reactions together and canceling electrons results in \[6 H^+ + I^– + 3 MnO_4^2– \rightarrow IO_3^– + 3 MnO_2 + 3 H_2O\] but this is writing as if in acid. We can add six \(OH^-\) ions to each side \[\cancel{6OH^- + 6 H^+} + I^– + 3 MnO_4^2– \rightarrow IO_3^– + 3 MnO_2 + \cancel{3 H_2O} + 6OH^- \label{Ex3}\] This neutralizes the hydronium ions on the reactant side of the equation to generate water (via \(H^+ + OH^- \rightarrow H_2O\)), which is partially canceled by the water in the products. Equation \(\ref{Ex3}\) is the balanced redox reaction in basic solution. \[3 H_2O + I^– + 3 MnO_4^{2–} \rightarrow IO_3^– + 3MnO_2 + 6 OH^– \] On the left, the six added OH– ions are combined with the six \(H^+\) ions of the acid-balanced equation to make 6 H2O. Three of these cancel with the 3 H2O on the right in the acid-balanced equation. Thus, we have a net of 3 H2O on the left in the base-balanced equation. All six OH– ions added on the right appear in the net redox reaction in base. Use the ion-electron method to complete and balance the following skeletal redox equations, occurring in either acidic or basic aqueous solution, as indicated. Identify the oxidation and reduction half reactions in each case. In acidic aqueous solution: \(Cu + NO3– \rightarrow Cu^{2+} + N_2O_4\) In acidic aqueous solution: \(XeO_3 + BrO_3^– \rightarrow Xe + BrO_4^-\) In acidic aqueous solution: \(MnO_4^– + CH_3OH \rightarrow Mn^{2+} + HCO_2H\) In acidic aqueous solution: \(Cr_2O_7^{2–} + I_2 \rightarrow Cr^{3+} + IO_3^–\) In basic aqueous solution: \(Pb(OH)_4^{2–} + ClO^– \rightarrow PbO_2 + Cl^–\) In basic aqueous solution: \(SO_2 + MnO_4^– \rightarrow SO_4^{2–} + MnO_2\)

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