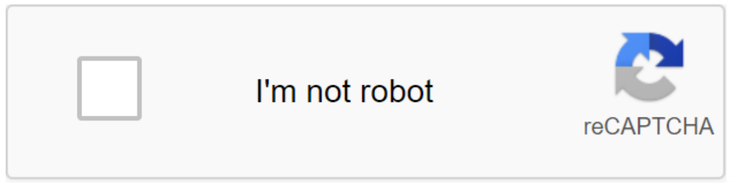

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Balancing redox equations worksheet

Determine the amount of oxidation of the elements in each of the following compounds: a. H2CO3 b. N2 c. (OH)42- d. NO2- e. LiH F. Fe3O4 Tip Identify the species oxidized and decreases in each of the following reactions: a. Cr' Sn4 Cr3 - Sn2 b. 3 Hg2 2 Fe (s) 3 Hg2 2 Fe3 c. 2 As (s) 3 Cl2 (g) 2 AsCl3 Hint Would You Use an Oxidant Agent or Reducing Agent to Make the Next Reactions Occur? a. ClO3- ClO2 b. SO42- C. Mn2 MnO2 d. yn'Cl2 Tip Write balance equations for the following redox reactions: a. NaBr - Cl2 NaCl - Br2 b. Fe2O3 - CO Fe - CO2 in Acid Solution C. CO - I2O5 CO2 - I2 in Basic Tip Tip Write Balanced Equations for the following reactions: Tip A. Cr(OH)3 - Br2 CrO42- Br- in base solution b. O2 - Sb H2O2 - SbO2 - in basic solution Tip c. HCOOH - MnO4- CO2 - Mn2 in sour solution d. ClO2- ClO2 - Cl- in sour solution Tip Write a balanced half of reactions NiO2 - 2 H2O - Fe Ni (OH)2 - Fe (OH)2 in base solution b. CO2 - 2 NH2OH CO and N2 - 3 H2O in base solution c. 2 NH - H2O2 - 2 Fe2 2 Fe3 - 2 H2O in sour solution d. H' - 2 H2O - 2 MnO4 - 5 SO2 2 Mn2 - 5 HSO4 - in the acid solution Hint Oxidation-reduction reactions, also called redox reactions, suggest the transmission of electrons from one species to another. These types of reactions are at the heart of energy production devices such as batteries and fuel cells. They are also involved in many of the electrochemical processes by which we obtain useful materials. The reaction in which one species transmits electrons to another is called a reaction to oxidation reduction, also called the radish reaction. For example, we might think of a metal iron reaction with chlorine gas, to form iron ion chloride (III) as a pure transfer of six electrons from two iron atoms into three chlorine molecules: Table No (PageIndex{1}): Redox Reaction Sample Multiplier Half Reaction Direction Electrons REDOX Reaction 2 (Fe rightarrow Fe'3'3') electrons on Oxidation 3 (Cl_2 0 2e - rightarrow 2Cl)) electrons pulled out reducing the addition of scaled reactionary and scaled products leads to 2Fe0 3 Cl_2 x 0 x 6 e-6 Right- 2Fe 3 (cancel 6e- 2Cl - and the cancellation of electrons leads to the final redox reacton, 2Fe (s) - 3Cl_2 (g) rightarrow 2Fe 3 Fe pushes electrons and Cl2 pulls electrons, thus ensuring the transmission of electrons. one for oxidation and one for contraction. Note that in the reaction of half of the oxidation, the electrons appear on the right, and in the reaction of half the reduction they appear on the left. Every half multiplied by the factor, so that the number of electrons produced by oxidation is equal to the number consumed by the reduction. Oxidation and reduction are always associated with the transmission of electrons. Thus,

there is never oxidation without contraction and vice versa in the reaction of redox. Oxidation of something has to lead to something else to be reduced, and vice versa. Thus, the substance oxidized is considered an agent of the reduction of another substance, and the substance is reduced is seen as an agent of oxidation of another substance. This leads to the following definitions: an oxidizing agent (oxidizer) - a substance that causes an oxidized other substance and itself decreases. Reducing the agent (reducer) - a substance that causes the decline of another substance and oxidizes itself. In these terms, all redox reactions take on the general form of Ox_1 Red_2 rightarrow Red_1 Ox_2 In this general representation, Ox1 and Ox2 will oxidize substances (oxidizers), and Red1 and Red2 reducing substances (reducers). So when Ox1 reacts with Red2, it becomes its diminished species, Red1, at the same time Red2 becomes its oxidized species, Ox2. The right Ox_1 process is a Red_1 that can require, say, electrons. So we could write this as a reduction in half of the reaction (Ox_1) no- (rightarrow Red_1 label half1) Also, the process (Red_2 rightarrow Ox_2) is oxidation that can require, say, m electrons. So we could write this as an oxidation of half the reaction (Red_2 'rightarrow Ox_2' - label half2 To write a balanced redox reaction, we want to put these two halves of the reaction (Equations (ref half1) and ref half2) together so that there are no pure electrons to show on either side of the overall reaction equation. Assuming that (n eq m), we will need to multiply half the reaction reduction to factor m and oxidation half the reaction factor n, so that when you add the total number of electrons on both sides cancel. So We will bring our two halves of the reaction together as follows: Table (PageIndex{2}): General Redox Reaction Balancing Multiplier Half Reaction Direction Electrons REDOX Reaction (m Red_2) (Ox_1 Red_1) rightarrow Ox_2 - me'-) electrons pulled oxidation Adding scalable reactionary and scalable alogn products with electrons lifting leads to a balanced reaction of the redox nOx_2 mRed_1 (mOx_1 nRed_2) reaction to H (m) and second to (n)) , we have balanced the overall reaction of redox in terms of transmission of nm electrons, but these electrons do not appear in the final balanced equation. This should always be the case with a balanced redox reaction. Redox also causes a change in the number of gearbox oxidation and oxidative. In the one element in the form experiences a decrease in its number of oxidation, while in oxidation the opposite occurs. This is evidenced by the expanind Table (PageIndex{1}). Table (PageIndex{3}): Redox Reaction Sample Multiplier Half Reaction Change in Oxidative Reaction Number 2 (Fe'0 Rightarrow Fe » q q 3e) Fe oxidation number increases (0 rightarrow 3) Oxidation 3 (Oxidation 3 (Cl_2 0 x 2e) rightarrow 2Cl-) The number of Cl oxidation decreases (0 rightarrow -1) The reduction of two halve reactions in the table (PageIndex{1}) also illustrates another important feature of balancing redox reactions. Note that each half of the reaction has a balance in both the number of atoms of each species and the overall charge on each side. For example, in the oxidative case Fe0 (rightarrow Fe) 3e- we have one iron atom on each side, but also zero charging on the left is balanced by a sum of 3 (3)) to the right of the Fe3 ion and three negative electrons. Similarly, when the Cl_2 2e- rightarrow 2Cl is reduced, the charge -2 to the left of the two electrons is balanced by a charge (2) (-1) to the right of the two chloride ions. In each half of the reaction and each common redox equation there should be both mass balance and charge balance. For each of the following, separate the skeletal (unbalanced) equation into two halves of the reaction. For each half of the reaction, balance the elements (mass balance) and then add the electrons to the right or left side to make a clean charge balance. Determine which half of the reaction is oxidation and which is the contraction. Then multiply each half of the reaction by the corresponding factor so that the two multiplied halve reactions are merged together to make a balanced redox equation. (Hg_2 x S_2O_3 2) (Al-Kru- 3) S_4O_6 (right) Kroo 2) (Au-3) (I - rightarrow au I_2)) There are two main methods of balancing redox equations: the method of ion-electronic oxidation method. The latter is easier to use with redox reactions in an aqueous solution and if necessary can be adapted to many situations that are not in aqueous solutions. Our main interest will be in the aqueous-solution redox; so we will use the ion-electronic method. One of the main advantages of this method is that it makes it completely unnecessary to assign individual oxidation numbers. To balance the redox equation with an ion-electronic method, take the following steps in this sequence: Divide skeletal equations into two halves of reactions. One half of the reaction will be reduced and the other will be oxidation. At this stage there is no need to determine what exactly. Balance each half of the reaction separately. Balance of atoms on each side of half of the reaction when examined. If the reaction occurs in an acidic environment, you can add H2O and/or H to balance oxygen and/or hydrogen. Going into a base environment, you can add H2O and/or OH-balance oxygen and/or hydrogen. Do not add any other new species (e.g. O2, H2) if it is already part of the skeletal half reaction. Balance the clean charge through each half of the reaction by adding electrons to the side with a more positive clean ion charge. If this process electrons are added on the left side of half the reaction, half of the reaction is a reduction. If electrons are added to the right side, half the oxidation reaction. (If you add electrons in the same direction in both halves, something is wrong!) Multiply both halves of the reaction by the corresponding factors of the entire number, so that the number of electrons is the same in both half of the reactions and will cancel when the two are added together. Add two multiplied half reactions together to get a common redox equation. Check the balance. Electrons should not appear in the general redox equation. Not only should there be an element-by-element balance throughout the equation, but the net charge (the amount of both ion charges and electronic charges) on both sides of the equation should be equal. Please note that this procedure does not involve assigning oxidation numbers. However, if the oxidation numbers are assigned to a balanced equation, there will always be a reduction in the acidity of an element, and oxidation implies an increase in the oxidation status of an element. The following examples illustrate the ion-electronic procedure, starting with the skeletal equation in the acidic or base solution. Balance (NO_3) rightarrow HNO_2 Fe 3) in acidic aquistric solution. The fact that this reaction occurs in an acidic aquision suggests that water or (H) may be involved in the reaction. This is not a requirement from the issue, of course. Species protonation NO_3) further shows HNO_2 this. Table (PageIndex{4}: Redox Reaction Sample Animation Half Reaction Change in Oxidative Reaction Number 2 (Fe)2 (Right Farrow Fe) 2 Right Oxidation 3 (2e) 3 H NO_3 rightarrow HNO_2 H_2O) N oxidation number decreases (N oxidation number decreases (N oxidation number decreases N oxidation number decreases (N oxidation number decreases (N oxidation number decreases N oxidation number decreases (N oxidation number decreases (N oxidation number decreases (N oxidation number decreases (No 5 rightarrow 3) Reducing adding these half reactions together and canceling electrons results in 2 Fe 3 H NO_3 rightarrow 2 fe th HNO_2 H_2O to make oxygen balance in response to the reduction of half HNO_2 NO_3 (H_2O), we added to the right, and then added 3 H to make a hydrogen balance. These are the only acceptable species that are used in acidic environments. Balance (NiO_2 Cd (rightarrow Ni (OH) 2 - Cd (OH) in the basic aqueous solution. Of course. Table (PageIndex{5}: Redox Reaction Sample Multiplier Half Reaction Change in Oxidative Reaction Number 1 (2 OH) - Cd (rightarrow Cd (OH NiO_2 H_2O) 2)) Ni Oxidation Number Decreases (5rightarrow 3) Reduction Adding Reduction these half reactions together and the cancellation of electrons leads to NiO_2 Cd H_2O rightarrow ni (OH) 2 Cd (OH) because it is in the base, we can only add H2O and/or OH- to make oxygen and hydrogen. The need to add OH-in Cd/Cd (OH)2 half of the reaction is simple. In the semi reaction of NiO2/Ni, think of H2O as an acid neutralizing the basic NiO2. So we add two H2Os on the left to neutralize two NiO2 O2 ions, and then we add two OH-to-right to complete the balance. Balancing oxygen and hydrogen in basic redox reactions can sometimes be difficult because both OH- and H2O contain both elements. The trick to getting around this is to balance any troublesome semi-reaction or whole redox reaction at first, as if it were in acid using H and H2O. The acid-balanced equation is then converted into its form in the base environment, adding the same amount of OH to both sides of the equation, which would be necessary to neutralize any H in the acid-balanced equation. Combine H and OH pairs to become H2O; i.e., (HH - OHH- right-H_2O). The following example shows this method for the redox reaction to be balanced in the base. Balance (I) - MnO_4 2- rightarrow IO_3 and MnO_2) in the basic aqueous solution. First, we will balance this in acid, and then neutralize any reaction to redox in basic conditions. Table :(PageIndex{6}: Redox Reaction Sample Sample Half Reaction Change in Oxidative Reaction Number 1 (3 H_2O IO_3) I'm Oxidizing Number Increases (-1 Rightarrow 5) Oxidation 1 (2e - 2 H_2O NiO_2 Rightarrow Ni (OH)2 2 2 OH. -) The number of Ni oxidation decreases (4 rightarrow 2) Reducing Adding these half reactions together and cancelling electrons resulting in 6 H - 3 MnO_42- rightarrow IO_3 3 MnO_2 3 H_2O but this is a writing as if acid is in acid. We can add six ions in each direction to cancel 6 HZ, I 3 MnO_4 x 2 - right IO_3 H_2O 3 MnO_2 This neutralizes hydronia ions on the reactionary side of the equation for water generation (through H_2O H and OH) products), which is partially offset in water. Equation (refEx3) is a balanced redox reaction in a basic solution. (3 H_2O) - 3 MnO_4 x2-2 right IO_3 - 3MnO_2 - 6 OH left, six added Ions OH- combined with six ions acid-balanced make 6 H2O. Three of them are canceled with 3 H2O on the right in the acid-balanced equation. So we have a 3 H2O network left in the basic balanced equation. All six OH-ions added to the right appear in the pure redox reaction in the base. Use the ion-electronic method to complete and balance the following skeletal redox equations occurring in acidic or base axvoid solution as indicated. Identify oxidation and reduce half of the reaction in each case. In acidic aquier solution: (Cu no3- rightarrow Cu2) N_2O_4 in acidic aquisea solution: (CH_3OH MnO_4 BrO_4 BrO_3 XeO_3 HCO_2H) In acidic aquier solution: (Cr_2O_7 I_2 1 2- IO_3 In basic aquision solution: PbO_2 (Pb (OH) : (SO_2 MnO_4 - right SO_4 and MnO_2) MnO_2)) balancing redox equations worksheet answers. balancing redox equations worksheet answers chemistry if8766. balancing redox equations worksheet answer key. balancing redox equations worksheet page 93. balancing redox reactions worksheet pdf. balancing redox equations worksheet a level. rules for balancing redox equations worksheet. balancing redox equations oxidation number method worksheet