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Equation Examples [1-5] Solved Problems Hess's law states that the total enthalpy change of a reaction equals the sum of all the enthalpy changes occurring in each step of the reaction. In other words, net enthalpy is independent of the number of steps the reaction takes to complete. Hess's law is a consequence of the conservation of energy [1-4]. This law is essential because it establishes enthalpy as a state function. It allows us to calculate the overall enthalpy change by adding up the enthalpy changes for each step until the product is formed. For a reaction to obey Hess's law, all the steps of the reaction must start and end at constant temperature and pressure. Hess Law This law is named after Russian chemist Germain Hess, who published it in 1840. According to Hess's law, enthalpy changes arising from the different steps are additive. Therefore, the overall enthalpy change of the reaction is the sum of all the enthalpy changes [1-5]. $\Delta Hrxn = \Sigma \Delta Hn$ $\Delta Hrxn$: Enthalpy change of the reaction ΔHn: Enthalpy change for each step Example 1. Formation of carbon dioxide (CO2) from graphite (C) and oxygen (O2) C (s, graphite) + O2 (g) → CO2 (g) The above reaction occurs in two steps. Step 1: Graphite reacts with oxygen to form carbon monoxide (CO) gas, releasing -110 kJ of energy per mole of CO. C (s, graphite) + ½ O2 (g) → CO (g) ΔH = -110.5 kJ/mol Step 2: Carbon monoxide reacts with oxygen forming carbon dioxide, releasing 283 kJ of energy per mole of CO2. CO (g) + ½ O2 (g) → CO2 (g) ΔH = -283.0 kJ/mol Combining the two equations, we get the following equation: C (s, graphite) + ½ O2 (g) + CO (g) + ½ O2 (g) → CO2 (g) (or, C (s, graphite) + O2 (g) → CO2 (g)) Which is the equation for formation of CO2. According to Hess's law, we can combine the enthalpies to get the net enthalpy change or the enthalpy of formation of CO2 gas. $\Delta Hrxn = \Delta Hfo$ = -110.5 kJ/mol - 283.0 kJ/mol = -393.5 kJ/mol The enthalpy change for the above reaction is represented in the diagram below. Hess Law Example Example 2. Formation of sulfur trioxide (SO3) from sulfur (S) and oxygen (O2). S (s) + 3/2 O2 (g) → SO3 (g) The above reaction occurs in two steps. Step 1: Sulfur reacts with oxygen to form sulfur dioxide (SO2), releasing 297 kJ of energy per mole of SO2 S (s, graphite) + O2 (g) → SO2 (g) ΔH = -297 kJ/mol Step 2: Sulfur dioxide reacts with oxygen to produce sulfur trioxide and 98 KJ of energy per mole of SO3 SO2 (g) + ½ O2 (g) → SO3 (g) ΔH = -98 kJ/mol Combining the two equations S (s, graphite) + O2 (g) + SO2 (g) + ½ O2 (g) → SO3 (g) Or, S (s) + 3/2 O2 (g) → SO3 (g) Which is the equation for formation of SO3 gas. Using Hess's law to combine the two enthalpy changes $\Delta Hrxn = \Delta Hfo$ = -297 kJ/mol - 98 kJ/mol = -395 kJ/mol Problem 1: Find the net enthalpy change ($\Delta Hrxn$) of the reaction below. CS2 (l) + 3 O2 (g) → CO2 (g) + 2 SO2 (g) The reactions in each step and their corresponding enthalpy change are given below. C (s) + O2 (g) → CO2 (g) ΔH = -393.5 kJ/mol (i) S (s) + O2 (g) → SO2 (g) ΔH = -295 kJ/mol (ii) C (s) + 2S (s) → CS2 (l) ΔH = +90 kJ/mol (iii) Solution 1. We leave reaction (i) as it is since CO2 is a product in the primary reaction. 2. We multiply reaction (ii) by 2 since we want 2 moles of CO2. The enthalpy change will also be multiplied by 2. 2 S (s) + 2 O2 (g) → 2 SO2 (g) ΔH = -590 kJ/mol (iv) 3. We flip reaction (iii) since we want CS2 to be a reactant. CS2 (l) → C (s) + 2S (s) ΔH = -90 kJ/mol (v) Adding all the reactions together C (s) + O2 (g) + 2 S (s) + 2 O2 (g) + CS2 (l) → CO2 (g) + 2 SO2 (g) + C (s) + 2 S (s) Or, CS2 (l) + 3 O2 (g) → CO2 (g) + 2 SO2 (g) ΔH = -1073.5 kJ/mol Hess's Law helps calculate enthalpy by adding up steps with known enthalpy values. You can reverse or multiply reactions to find the enthalpy change in a chemical reaction. Hess's Law can also find entropy and Gibb's energy along with enthalpy changes. Hess's Law, also known as "Hess's Law of Constant Heat Summation," states that the total enthalpy of a chemical reaction is the sum of the enthalpy changes for the steps of the reaction. Therefore, you can find enthalpy change by breaking a reaction into component steps that have known enthalpy values. This example problem demonstrates strategies for how to use Hess's Law to find the enthalpy change of a reaction using enthalpy data from similar reactions. What is the value of ΔH for the following reaction? CS2(l) + 3 O2(g) → CO2(g) + 2 SO2(g) Given: C(s) + O2(g) → CO2(g), ΔH= -393.5 kJ/molS(s) + O2(g) → SO2(g), ΔH = -296.8 kJ/molC(s) + 2 S(s) → CS2(l), ΔH= 87.9 kJ/mol Hess's Law says the total enthalpy change does not rely on the path taken from beginning to end. Enthalpy can be calculated in one grand step or multiple smaller steps. To solve this type of problem, organize the given chemical reactions where the total effect yields the reaction needed. There are a few rules that you must follow when manipulating a reaction. The reaction can be reversed. This will change the sign of ΔH.The reaction can be multiplied by a constant. The value of ΔH must be multiplied by the same constant.Any combination of the first two rules may be used. Finding a correct path is different for each Hess's Law problem and may require some trial and error. A good place to start is to find one of the reactants or products where there is only one mole in the reaction. You need one CO2, and the first reaction has one CO2 on the product side. C(s) + O2(g) → CO2(g), ΔH = -393.5 kJ/mol This gives you the CO2 you need on the product side and one of the O2 moles you need on the reactant side. To get two more O2 moles, use the second equation and multiply it by two. Remember to multiply the ΔH by two as well. 2 S(s) + 2 O2(g) → 2 SO2(g), ΔH = 2(-326.8 kJ/mol) Now you have two extra S's and one extra C molecule on the reactant side that you don't need. The third reaction also has two S's and one C on the reactant side. Reverse this reaction to bring the molecules to the product side. Remember to change the sign on ΔH. CS2(l) → C(s) + 2 S(s), ΔH = -87.9 kJ/mol When all three reactions are added, the extra two sulfur and one extra carbon atoms are canceled out, leaving the target reaction. All that remains is adding up the values of ΔH. ΔH = -393.5 kJ/mol + 2(-296.8 kJ/mol) + (-87.9 kJ/mol)ΔH = -393.5 kJ/mol - 593.6 kJ/mol - 87.9 kJ/molΔH = -1075.0 kJ/mol Answer: The change in enthalpy for the reaction is -1075.0 kJ/mol. Hess's Law takes its name from Russian chemist and physician Germain Hess. Hess investigated thermochemistry and published his law of thermochemistry in 1840. To apply Hess's Law, all of the component steps of a chemical reaction need to occur at the same temperature. Hess's Law may be used to calculate enthalpy and Gibb's energy in addition to enthalpy. In the previous post, we talked about how Hess law and how it can be used for measuring the enthalpy change (ΔH) for a reaction without making calorimetric measurements. In short, the Hess's law is based on the fact that enthalpy is a state function, and therefore, $\Delta Hrxn$ is the same whether the reaction takes place in one step or a series of steps. For example, the enthalpy change for the reaction 2A + B → 2D can be calculated using the enthalpies of the two reference reactions: When solving a Hess's law problem, remember your objective is to manipulate the reference equations such that when you add them up, the target equation is obtained. Follow these two rules when changing the reference equations: If the equation is multiplied by any factor, ΔH must be multiplied by the same factor. If the equation is reversed, the sign of ΔH must be changed. 1. Calculate the enthalpy for the oxidation of CO to CO2 using the enthalpy of reaction for the combustion of C to CO (ΔH = -221.0 kJ) and the enthalpy for the combustion of C to CO2 (ΔH = -393.5 kJ). 2CO(g) + O2(g) → 2CO2(g) ΔH = ? C(s) + O2(g) → CO2(g) ΔH = -393.5 kJ 2C(s) + O2(g) → 2CO(g) ΔH = -221.0 kJ 2. Calculate the enthalpy for the combustion reaction of sulfur to sulfur trioxide using the enthalpies of the two reactions shown below: 2S(s) + 3O2(g) → 2SO3(g) ΔH = ? S(s) + O2(g) → SO2(g) ΔH [SO3(g) → SO2(g) + O2(g)] ΔH = 198 kJ 3. Using the Hess's law, calculate ΔHo for the combustion reaction of butane: C4H8(g) + 6O2(g) → 4CO2(g) + 4H2O(l), ΔHo = ? Use the following reactions and given ΔH's: 1) 2H2(g) + O2(g) → 2H2O(g), ΔHo = -571 kJ 2) C4H8(g) + H2(g) → C4H10(g), ΔHo = -126 kJ 3) 2C4H10(g) + 13O2(g) → 8CO2(g) + 10H2O(l), ΔHo = -5754 kJ 4. Using the Hess's law and the enthalpies of the three combustion reactions below, calculate the enthalpy of the reaction producing methanol (CH3OH) from carbon monoxide and hydrogen gas. CO(g) + 2H2(g) → CH3OH(g), ΔH = ? 1) 2CO(g) + O2(g) → 2CO2(g), ΔH = -566 kJ 2) 2H2(g) + O2(g) → 2H2O(g), ΔHo = -571 kJ 3) 2CH3OH(g) + 3O2(g) → 2CO2(g) + 4H2O(g), ΔHo = -1430 kJ 5. Using the Hess's law and the enthalpies of the given reactions, calculate the enthalpy of the combustion reaction of CH3Cl: CH3Cl(g) + O2(g) → CO(g) + HCl(g) + H2O(l), ΔH = ? 1) CO(g) + 2H2(g) → CH3OH(g), ΔHo = 139 kJ 2) 2H2(g) + O2(g) → 2H2O(g), ΔHo = -571 kJ 3) CH3OH(g) + HCl(aq) → CH3Cl(g) + H2O(l), ΔHo = -28 kJ 6. Using the Hess's law and the enthalpies of the given reactions, calculate the enthalpy of the following oxidation reaction between CuO and HCl: 2CuO(s) + 4HCl(g) → 2CuCl(s) + Cl2(g) + 2H2O(g), ΔH = ? 1) CuO(s) + H2(g) → Cu(s) + H2O(g), ΔH = -85 kJ 2) 2Cu(s) + Cl2(g) → 2CuCl(s), ΔH = -274 kJ 3) H2(g) + Cl2(g) → 2HCl(g), ΔH = -184 kJ 8. Determine the enthalpy of formation of methane using the following data obtained from bomb calorimetry: C(graphite) + 2H2(g) → CH4(g) ΔH °f = ? CH4(g) + 2O2(g) → CO2(g) + 2H2O(l), ΔH ° = -891 kJ C(s) + O2(g) → CO2(g) ΔH = -393 kJ 2H2(g) + O2(g) → 2H2O(l) ΔH = -572 kJ Hess's law of constant heat summation, or Hess's law for short is a relationship describing the enthalpy change of a reaction. The total enthalpy change of a reaction is the sum of total enthalpies for each step of the reaction and is independent of the order of the steps. Basically, calculate the total enthalpy by breaking a reaction down to simple component steps of known enthalpy values. This Hess's Law example problem shows how to manipulate reactions and their enthalpy values to find the total change of enthalpy of a reaction.First, there are a couple notes to keep straight before beginning.If a reaction is reversed, the sign of the change in enthalpy (ΔH) changes.For example: the reaction C(s) + O2(g) → CO2(g) has an ΔH of -393.5 kJ/mol. The reverse reaction CO2(g) → C(s) + O2(g) has a ΔH of +393.5 kJ/mol.If a reaction is multiplied by a constant, the change in enthalpy is changed by the same constant.Example, for the previous reaction, if three times the reactants are allowed to react, ΔH is changed by three times.If ΔH is positive, the reaction is endothermic. If ΔH is negative, the reaction is exothermic.Question: Find the enthalpy change for the reactionCS2(l) + 3 O2(g) → CO2(g) + 2 SO2(g)when:C(s) + O2(g) → CO2(g); ΔH = -393.5 kJ/molS(s) + O2(g) → SO2(g); ΔH = -296.8 kJ/molC(s) + 2 S(s) → CS2(l); ΔH = 87.9 kJ/molSolution: Hess's Law problems can take a little trial and error to get started. One of the best places to begin is with a reaction with only one mole of reactant or product in the reaction.Our reaction needs one CO2 in the product and the first reaction also has one CO2 product.C(s) + O2(g) → CO2(g) ΔH = -393.5 kJ/molThis reaction gives us the CO2 needed on the product side and one of the O2 needed on the reactant side. The other two O2 can be found in the second reaction.S(s) + O2(g) → SO2(g) ΔH = -296.8 kJ/molSince only one O2 is in the reaction, multiply the reaction by two to get the second O2. This doubles the ΔH value.2 S(s) + 2 O2(g) → 2 SO2(g) ΔH = -593.6 kJ/molCombining these equations gives2 S(s) + C(s) + 3 O2(g) → CO2(g) + SO2(g)The enthalpy change is the sum of the two reactions: ΔH = -393.5 kJ/mol + -593.6 kJ/mol = -987.1 kJ/molThis equation has the product side needed in the problem but contains an extra two S and one C atom on the reactant side. Fortunately, the third equation has the same atoms. If the reaction is reversed, these atoms are on the product side. When the reaction is reversed, the sign of the change in enthalpy is reversed.CS2(l) → C(s) + 2 S(s), ΔH = -87.9 kJ/molAdd these two reactions together and the extra S and C atoms cancel out. The remaining reaction is the reaction needed in the question. Since the reactions were added together, their ΔH values are added together.ΔH = -987.1 kJ/mol + 87.9 kJ/molΔH = -1075 kJ/molAnswer: The change in enthalpy for the reactionCS2(l) + 3 O2(g) → CO2(g) + 2 SO2(g) ΔH = -1075 kJ/mol.Hess's law requires reassembling the component reactions until the total enthalpy is achieved. While Hess's Law applies to changes in enthalpy, this law can be used for other thermodynamic state equations such as Gibbs energy and entropy.Related Posts Problem #1: The heats of combustion of C(s, graphite), H2(g) and CH4(g) at 298 K and 1 atm are respectively −393.50 kJ/mol, −285.83 kJ/mol and −890.36 kJ/mol. What is the enthalpy of formation for CH4? Solution 1) The three combustion reactions are: C(s, gr) + O2(g) ---> CO2(g) ΔH = −393.50 kJ H2(g) + 1/2O2(g) ---> H2O(l) ΔH = −285.83 kJ CH4(g) + 2O2(g) ---> CO2(g) + 2H2O(l) ΔH = −890.36 kJ The phrase 'enthalpy of formation' tells us this reaction is called for: C(s, gr) + 2H2(g) ---> CH4(g) 3) I saw this answer on an "answers" type website and decided to include it: This is a Hess's Law problem. If you multiply the first reaction by 1, the second by 2, and the third by negative 1 (write it backwards) then add together to give the reaction you're looking for. So, the enthalpy of the reaction you're solving for is equal to 1(−393.50) + 2(−285.83) + (−1)(−890.36). I'll let you finish it. In the work, here is how the ChemTeam usually solves these problems: C(s, gr) + O2(g) ---> CO2(g) ΔH = −393.50 kJ 2H2(g) + O2(g) ---> 2H2O(l) ΔH = −571.66 kJ When the three data equations are added together, duplicates on each side of the arrow will be removed. Here's before any removal: C(s, gr) + O2(g) + 2H2(g) + O2(g) + CO2(g) + 2H2O(l) ---> CO2(g) + 2H2O(l) + CH4(g) Here is the −74.80 kJ Problem #2: The standard heat of combustion of benzene is −3271 kJ/mol, for CO2 it is −393 kJ/mol, and for H2O, it is −286 kJ/mol. Calculate the standard heat of formation of benzene. (Note that the first bit of data is associated with the reactant (benzene) while the last two are associated with the products (CO2 produced when C combusts and H2O produced when H2 is combusted. The reason for this will be seen in solution #2.) Solution #1: 1) The term 'standard heat of formation' tells us that this equation is the desired target: 6C(s) + 3H2(g) ---> C6H6(ℓ) 2) Set up the three combustion equations C6H6 + 15 2O2 ---> 6CO2 + 3H2O ΔH = −3271 kJ C + O2 ---> CO2 ΔH = −394 kJ H2 + 1/2O2 ---> H2O ΔH = −286 kJ 3) Flip equation 1, multiply equations 2 and 3: 6CO2 + 3H2O ---> C6H6 + 15 2O2 ΔH = +3271 kJ 6C + 6O2 ---> 6CO2 ΔH = −2364 kJ 3H2 + 3O2 ---> 3H2O ΔH = −858 kJ 4) Add the three equations: 3271 + (−2364) + (−858) = +49 kJ/mol You may check the answer here. Problem #2: 1) Write the equation for the combustion of benzene can be calculated thusly: ΔH rxn = $\Sigma \Delta H$f, products − $\Sigma \Delta H$f, reactants Where x represents the standard heats of formation. 'comb' stands for combustion. 3) Inserting values and solving, we have: −3271 = [(6) (−394) + (3) (−286)] − [(1) (x) + (15 2) (0)] x = +49 kJ/mol The solution to the above depends on the fact that the combustion reactions for C and H2 are also the formation reactions for CO2 and H2O, with +394 kJ/mol the enthalpy of formation for H2O. Problem #3: The heat of combustion for the gases hydrogen, methane and ethane are −285.8, −890.4 and −1559.9 kJ/mol respectively at 298K. Calculate (at the same temperature) the heat of reaction for the following reaction: 2CH4(g) ---> C2H6(g) + H2(g) Solution: 1) The data given are these three reactions: H2 + 1/2O2 ---> H2O ΔH = −285.8 kJ CH4 + 2O2 ---> CO2 + 2H2O ΔH = −890.4 kJ C2H6 + 7/2O2 ---> 2CO2 + 3H2O ΔH = −1559.9 kJ 2) Manipulate the data equations as follows: a) flip (so hydrogen as a product) b) multiply by 2 (gives us 2CH4 on the reactant side) c) flip (puts C2H6 as a product) 3) Adding the three equations yields the answer to the problem, +64.9 kJ. Problem #4: What is the standard enthalpy of reaction for the reduction of iron(II) oxide by carbon monoxide? FeO(s) + CO(g) ---> Fe(s) + CO2(g) given the following information: 3Fe2O3(s) + CO(g) ---> 2Fe3O4(s) + CO2(g) ΔH = −48.26 kJ Fe2O3(s) + 3CO(g) ---> 2Fe(s) + 3CO2(g) ΔH = −23.44 kJ Fe3O4(s) + CO(g) ---> 3FeO(s) + CO2(g) ---> ΔH = +21.79 kJ Solution: 1) Changes to be made to the data equations: a) reverse equation 1 (this puts Fe3O4 on opposite side to compensate for switching equation 3) b) multiply equation 2 by 3 (this will give 3Fe2O3, allowing it to cancel) c) reverse equation 3 and multiply it by two (this puts FeO on the reactant side and gives us 2Fe3O4 to cancel) Please note that no attention was paid to CO and CO2. If everything else is done correctly, they should fall into line. 2) The three data equations with the changes applied: 2Fe3O4(s) + CO2(g) ---> 3Fe2O3(s) + CO(g) ΔH = +48.26 kJ 3Fe2O3(s) + 9CO(g) ---> 6Fe(s) + 9CO2(g) ΔH = −70.32 kJ 6FeO(s) + 2CO2(g) ---> 2Fe3O4(s) + 2CO(g) ΔH = −43.58 kJ 3) Adding the three equations together gives: 6FeO(s) + 6CO(g) ---> 6Fe(s) + 6CO2(g) ΔH = −65.64 kJ 2Fe3O4(s) + CO2(g) ---> 3Fe2O3(s) + CO(g) ΔH = +48.26 kJ b) Dividing through by six gives the final answer: FeO(s) + CO(g) ---> Fe(s) + CO2(g) ΔH = −10.94 kJ Problem #5: Determine the enthalpy of benzene can be calculated thusly: ΔH rxn = 3271 kJ The solution to the above depends on the fact that the combustion reactions for C and H2 are also the formation reactions for CO2 and H2O, with +394 kJ/mol the enthalpy of formation for H2O. Problem #3: The heat of combustion for the gases hydrogen, methane and ethane are... 2Fe3O4(s) + CO2(g) given the following data: Fe2O3(s) + 3CO(g) ---> 2Fe(s) + 3CO2(g) ΔH = −23.44 kJ Fe3O4 + CO(g) ---> 3FeO + CO2(g) ΔH = +21.79 kJ Fe2O3(s) + CO(g) ---> 2Fe3O4(s) + CO2(g) ΔH = −10.94 kJ Solution: 1) Apply the following changes to the data equations: a) multiply first equation by 3 (to give us 3Fe2O3) b) multiply through by 2 (to put 2Fe3O4 on the product side) c) multiply third equation by 6 (to cancel Fe and FeO) Note that I have ignored the CO and CO2. If everything works out, the right amounts will be there. 2) The result: 3Fe2O3(s) + 9CO(g) ---> 6Fe(s) + 9CO2(g) ΔH = −70.32 kJ 6FeO(s) + 2CO2(g) ---> 2Fe3O4 + 2CO(g) ΔH = −43.58 kJ 6Fe(s) + 6CO(g) ---> 6Fe(s) + 6CO2(g) ΔH = −65.64 kJ Add the three equations and the three equations and their enthalpy. It had enthalpy values which were not the final answers. Be aware of this practice (one with which I disagree). It is done to guard against someone finding the solved problem on the Internet with the correct answer. Problem #6: Iron metal can be produced in a blast furnace through a complex series of reactions involving reduction of iron(II) oxide with carbon monoxide. iron(II) oxide + carbon monoxide ---> iron + carbon dioxide Use the equations below to calculate ΔH for the overall equation. (a) 3Fe2O3(s) + CO(g) ---> 2Fe3O4(s) + CO2(g) ΔH = −48.26 kJ (b) Fe(s) + CO2(g) ---> FeO(s) + CO(g) ΔH = +10.94 kJ (c) Fe2O3(s) + 3CO(g) ---> 2Fe(s) + 3CO2(g) ΔH = +21.79 kJ Solution: 1) Let's get a balanced equation for our target reaction: Fe2O3 + 3CO ---> 2Fe + 3CO2 2) Rearrange the three data equations so that, when added, they give the target equation: a) leave untouched b) flip, multiply by 6 c) multiply by 2 3) This results in: (a) 3Fe2O3(s) + CO(g) ---> 2Fe3O4(s) + CO2(g) ΔH = −48.26 kJ (b) 6FeO(s) + 6CO(g) ---> 6Fe(s) + 6CO2(g) ΔH = −65.64 kJ (c) 2Fe3O4(s) + 6FeO(s) + 2CO2(g) ---> 2Fe3O4(s) + 2CO(g) ΔH = +43.58 kJ 4) When the three equations are added together, this results in: 3Fe2O3(s) + 9CO(g) ---> 6Fe(s) + 9CO2(g) ΔH = −48.26 + (−65.64) + 43.58 = −70.32 kJ 5) To get the final answer, divide everything by 3: Fe2O3(s) + 3CO(g) ---> 2Fe(s) + 3CO2(g) ΔH = −23.44 kJ Problem #7: From the following data: N2 + 3/2O2 ---> N2O3 ΔH = 83.7 kJ N2 + O2 ---> 2NO ΔH = 180.4 kJ 1/2N2 + O2 ---> NO2 ΔH = 33.2 kJ Calculate the enthalpy change for the reaction: N2O3 ---> NO + NO2 Solution: 1) What we need to flip the first data equation. We do this to put N2O3 on the reactant side. b) We will divide the second data equation by 2. We do this to get NO (which is already a product, it is where we want it) rather than 2NO. c) The third data equation will be left alone. We have NO2 on the product side and it's the right coefficient. 2) When I apply the above changes, I get this: N2O3 ---> N2 + 3/2O2 ΔH = −83.7 kJ 1/2N2 + 1/2O2 ---> NO ΔH = 90.2 kJ 1/2N2 + O2 ---> NO2 ΔH = 33.2 kJ 3) When you add the three above equations, you will recover the target equation. The N2 will cancel as well as the 3/2O2. Add the three enthalpies for the answer. Problem #8: Determine the standard enthalpy of formation for NO: 1/2N2(g) + 1/2O2(g) ---> NO(g) using the following three data equations: N2(g) + 3H2(g) ---> 2NH3(g) ΔH = −91.8 kJ 4NH3(g) + 5O2(g) ---> 4NO(g) + 6H2O(g) ΔH = −906.2 kJ 2H2(g) + O2(g) ---> 2H2O(g) ΔH = −483.6 kJ Solution: In the second equation, the 4NO is going to drive what I do to solve this problem. At the very end, I'm going to divide by 4. First equation: multiply it by 2. That is going to get me 4NH3 which will cancel with the 4NH3 in equation. Second equation: leave it alone Third equation: flip it and multiply by 6. That gets H2O on the left-hand side and the 6 will cancel the 6H2O in the second equation. Notice that I'm not paying attention to the N2 or O2. That's because, if I do everything else right, they will come out correctly as well. Here are the equations with the changes: 2N2(g) + 6H2(g) ---> 4NH3(g) ΔH = −183.6 kJ 4NH3(g) + 5O2(g) ---> 4NO(g) + 6H2O(g) ΔH = −906.2 kJ 6H2O(g) ---> 6H2(g) + 3O2(g) ΔH = +1450.8 kJ When you add the three equations together, the 6H2 cancels as well as the 4NH3 and 3O2 (of the four). Here is what results: 2N2(g) + 2O2(g) ---> 4NO(g)ΔH = +361 kJ Last step: divide the equation by 4 to get the target equation. Divide the enthalpy by 4 to obtain +90.25 kJ Is this the correct answer? Of course it is! By the way, you could also start by dividing the second equation by 4. I will leave you to contemplate what to do with the first and third data equations after dividing equation 2 by 4. Problem #9: Use Hess's Law to determine ΔH for the reaction: NO + O ---> NO2 given the following reactions O2 ---> 2O ΔH = +495 kJ 2O3 ---> 3O2 ΔH = −142.3 kJ NO + O3 ---> NO2 + O2 ΔH = −199 kJ Solution: O ---> 1/2O2 ΔH = −247.5 kJ (flipped, gets O as a reactant, divide by 2, gets one O3 ΔH = −71.15 kJ (flipped, divide by 2, sets up for O3 to cancel) NO + O3 ---> NO2 + O2 ΔH = −199) = −375.35 kJ Rounding off to −375 kJ seems like the best choice for the final answer. Problem #10: Using Hess' Law, determine the ΔH of the following reaction: N2(g) + 2O2(g) ---> 2NO2(g) Given the following equations N2(g) + 2H2(g) ---> 2NH3(g) ΔH = −115 kJ 2NH3(g) + 4H2O(r) ---> 2N2(g) + 7H2(g) ΔH = −142.5 kJ H2O(r) ---> H2(g) + 1/2O2(g) ΔH = −43.7 kJ Solution: Use the first two data equations without modifying them. For the third data equation, reverse it and multiply it by four. The H2 will cancel as will the NH3 and H2O. The final answer is −82.7 kJ. Bonus Problem: The standard molar enthalpy of formation, ΔH fo , of diborane cannot be determined directly because the compound cannot be prepared by reaction of boron and hydrogen. However, the value can be calculated. Calculate the standard enthalpy of formation of gaseous diborane (B2H6) using the following thermochemical information: (a) 4B(s) + 3O2(g) ---> 2B2O3(s)ΔH° = −2509.1 kJ (b) 2H2(g) + O2(g) ---> 2H2O(l)ΔH° = −571.7 kJ (c) B2H6(g) + 3O2(g) ---> B2O3(s) + 3H2O(l)ΔH° = −2147.5 kJ Solution: 1) An important key is to know what equation we are aiming for. The answer is in the word 'formation:' 2B + 3H2 ---> B2H6 Remember that formation means forming one mole of our target substance. This means that a one MUST be in front of the B2H6 2) In order to get ΔH of formation, this is the equation we want by two equation (b) - multiply through by 3/2 equation (c) - flip 3) Why? equation (a) - this gives us 2B (from 4B) for our final equation equation (b) - this gives us 3H2 for our final equation equation (c) - this puts B2H6 on the right-hand side of the final equation 4) The above manipulations have consequences for the coefficients AND the ΔH° values. Rewrite equations (a), (b) and (c): a) 2B(s) + 3/2O2(g) ---> B2O3(s)ΔH° = −1254.55 kJ b) 3H2(g) + 3/2O2(g) ---> 3H2O(l)ΔH° = −857.55 kJ c) B2O3(s) + 3H2O(l) ---> B2H6(g) + 3O2(g)ΔH° = +2147.5 kJ 5) Add the three equations and the ΔH° values to get: 2B + 3H2 ---> B2H6 ΔH = +35.4 kJ Here is the same diborane question, with slightly different numbers: a) 4B(s) + 3O2(g) ---> 2B2O3(s)ΔH° = −2543.8 kJ b) 2H2(g) + O2(g) ---> 2H2O(g)ΔH° = −484.0 kJ c) B2H6(g) + 3O2(g) ---> B2O3(s) + 3H2O(l)ΔH° = −2032.9 kJ Calculate the standard enthalpy of formation of gaseous diborane (B2H6). Just remember: With all Hess's Law (of heat summation) problems, the chemical reactions summarize the main concepts of thermochemistry. The key to these problems is that whatever you do to the chemical reaction, you must do to the ΔH value. So, for example, if you reverse the equation, you must reverse the sign of ΔH. If you multiply the equation by 2, then you must multiply the ΔH value by 2. In this set of practice questions, we will summarize the main concepts of thermochemistry such as the relationship between internal energy, work and heat, exothermic and endothermic process, heat capacity, constant-volume calorimetry, the enthalpy, the standard enthalpies of formation and their use in determining the heat of the reaction, and the Hess's law. The links to the corresponding topics are given below: 1. What is the sign of ΔE system if energy flows from the surroundings into a chemical system? 2. A gas absorbs 62 kJ of heat and does 48 kJ of work. Calculate the change in internal energy of the system if it releases 415 kJ of heat and does 562 kJ of work on the surroundings. 4. Calculate ΔE for each of the following. a) q = −62 kJ, w = +89 kJ b) q = −48 kJ, w = −92 kJ, w = 0 kJ 5. The gas in a piston is heated up by absorbing 544 J of heat and expands performing 541 J of work on the surroundings. What is the change in internal energy? 6. A pump is held under 1.30 atm external pressure. How much work (in J) is required to expand its volume from 1.80 L to 3.50 L? 7. A balloon filled with 8.50 moles of nitrogen at 23.0 °C expands from 3.54 L to 5.64 L by increasing its temperature to 58.0 oC under constant pressure of 1.10 atm. Calculate q, w, and ΔE in kJ for the nitrogen gas in the balloon if its heat capacity is 29.15 J/°C · mol. 8. How much heat does it take to increase the temperature of a 540.6-g sample of Fe from 20.0 °C to 84.3 °C? The specific heat of iron = 0.450 J/g °C. 9. Calculate the specific heat capacity of a metal if a 17.0 g sample requires 481 J to change the temperature of the metal from 25.0 °C to 67.0 °C? 12. Calculate the energy of combustion for one mole of butane if burning a 0.367 g sample of butane (C4H10) has increased the temperature of a bomb calorimeter by 7.73 °C. The heat capacity of the bomb calorimeter is 2.36 kJ/°C. 13. How many joules of energy is required to melt 40.0 g of ice at 0 °C? The heat of fusion (ΔHfus) for ice is 334.0 J/g. 14. How many kJ of energy does it take to change 36.0 g of ice at -15.0 °C to water at 0. °C? The specific heat of ice is 2.10 J/g°C and the heat of fusion (ΔHfus) for ice is 334.0 J/g. Ignore the significant figures for this problem. 15. The endothermic reaction is given below: 2CH3OH(l) + 3O2(g) → 4H2O(l) + 2CO2(g) What quantity of heat is released for each mole of water formed? 16. How much heat will be released if 44.8 g of SO2 is reacted with an excess of oxygen according to the following chemical equation? 2SO2(g) + O2(g) → 2SO3(g), ΔH = 198 kJ 17. What is ΔH° for the following reaction 2C6H6(l) + 15O2(g) → 12CO2(g) + 6H2O(l), ΔH = ? kJ if the consumption of 27.3 g of benzene (C6H6) produces 1144 kJ of heat? 18. Based on the heat of reaction for the chlorination of methane, how much heat will be released if 233.6 grams of hydrochloric acid are formed? CH4(g) + 3Cl2(g) → CHCl3(l) + 3HCl(g), ΔH° = −334 kJ 19. Calculate how many kJ of heat-energy will be released when 12.65 g of magnesium carbonate reacts with 650. mL of 0.400 M hydrochloric acid? MgCO3(s) + 2HCl(aq) → MgCl2(aq) + H2O(l) + CO2(g), ΔH° = −112 kJ 20. Using the standard heats of formation given below, calculate the heat of reaction for the following reaction, calculate the heat of reaction for the following reaction: 2C4H10(g) + 13O2(g) → 8CO2(g) + 10H2O(l), ΔH°rxn = −5755 kJ Calculate the molar enthalpy of formation of butane using the information given below: ΔHf° CO2(g) = −393.5 kJ/mol ΔHf° H2O(l) = −285.8 kJ/mol 22. Zinc is recovered from ZnS by first oxidizing it to ZnO. Calculate the enthalpy of this oxidation reaction using the data given below: ZnS(s) + 3O2(g) → ZnO(s) + 2SO2(g) ΔHf° for SO2(g) = −296.8 kJ/mol ΔHf° for ZnS(s) = −206.0 kJ/mol ΔHf° for ZnO(s) = −350.5 kJ/mol 23. Using the standard heats of formation given below, calculate the heat of reaction for... carbonate (BaCO3) and sulfuric acid (H2SO4). BaCO3(s) + H2SO4(aq) → BaSO4(s) + CO2(g) + H2O(l) ΔHf° for BaCO3(s) = −1473.2 kJ/mol ΔHf° for CO2(g) = −393.5 kJ/mol ΔHf° for H2O(l) = −285.8 kJ/mol ΔHf° for BaSO4(s) = −1213.0 kJ/mol ΔHf° for CO2(g) = −393.5 kJ/mol ΔHf° for H2O(l) = −285.8 kJ/mol 24. The heat of reaction for the nitrogen oxide and chlorine is −76 kJ: 2NO(g) + Cl2(g) → 2NOCl(g) What is the enthalpy change for the following reaction? 2NOCl(g) → 2NO(g) + Cl2(g) 25. The heat of reaction for the oxidation of hydrogen sulfide to sulfur dioxide accord according to the following reaction is -1036 kJ: 2H2S(g) + 3O2(g) → 2SO2(g) + 2H2O(g) What is the enthalpy change for the following reaction? 2SO2(g) + 2H2O(g) → 2H2S(g) + 3O2(g) 26. Consider the thermochemical equation for the combustion of nitromethane (CH3NO2): 2CH3NO2(g) + 3/2O2(g) → 2CO2(g) + 3H2O(g) + N2(g), ΔH °rxn = ? 27. Using the Hess's law, calculate ΔHo for the reaction: 2NO2 + Cl2(g) → 2NOCl(g) + O2(g), ΔHo = ? Use the following reactions and given ΔH's: 1) 2NO(g) + Cl2(g) → 2NOCl(g), ΔHo = −76 kJ 2) 2NO(g) + O2(g) → 2NO2(g), ΔHo = −114 kJ 28. Calculate the enthalpy for the oxidation of CO to CO2 using the enthalpy for the combustion of C to CO2 (ΔH = -393.5 kJ) 2CO(g) + O2(g) → 2CO2(g) ΔH = ? C(s) + O2(g) → CO2(g) ΔH = -393.5 kJ 2C(s) + O2(g) → 2CO(g) ΔH = -221.0 kJ 29. Calculate the enthalpy for the combustion reaction of butane: C4H8(g) + 6O2(g) → 4CO2(g) + 4H2O(l), ΔHo = ? Use the following reactions and given ΔH's: 1) 2H2(g) + O2(g) → 2H2O(g), ΔHo = -571 kJ 30. Using the Hess's law, calculate ΔHo for the combustion reaction of butane: C4H8(g) + 6O2(g) → 4CO2(g) + 4H2O(l), ΔHo = ? Use the following reactions and given ΔH's: 1) 2H2(g) + O2(g) → 2H2O(g), ΔHo = -571 kJ 2) C4H8(g) + H2(g) → C4H10(g), ΔHo = -126 kJ 3) 2C4H10(g) + 13O2(g) → 8CO2(g) + 10H2O(l), ΔHo = -5754 kJ 31. Using the Hess's law and the enthalpies of the three combustion reactions below, calculate the enthalpy of the reaction producing methanol (CH3OH) from carbon monoxide and hydrogen gas. CO(g) + 2H2(g) → CH3OH(g), ΔH = ? 1) 2CO(g) + O2(g) → 2CO2(g), ΔH = -566 kJ 2) 2H2(g) + O2(g) → 2H2O(g), ΔHo = -571 kJ 3) 2CH3OH(g) + 3O2(g) → 2CO2(g) + 4H2O(g), ΔHo = -1430 kJ 32. Using the Hess's law and the enthalpies of the given reactions, calculate the enthalpy of the combustion reaction of CH3Cl: CH3Cl(g) + O2(g) → CO(g) + HCl(g) + H2O(l), ΔH = ? 1) CO(g) + 2H2(g) → CH3OH(g), ΔHo = 139 kJ 2) 2H2(g) + O2(g) → 2H2O(g), ΔH = -571 kJ 3) CH3OH(g) + HCl(aq) → CH3Cl(g) + H2O(l), ΔH = -28 kJ 33. Using the Hess's law and the enthalpies of the given reactions, calculate the enthalpy of the following oxidation reaction between CuO and HCl: 2CuO(s) + 4HCl(g) → 2CuCl(s) + Cl2(g) + 2H2O(g), ΔH = ? 1) CuO(s) + H2(g) → Cu(s) + H2O(g), ΔH = -85 kJ 2) 2Cu(s) + Cl2(g) → 2CuCl(s), ΔH = -274 kJ 3) H2(g) + Cl2(g) → 2HCl(g), ΔH = -184 kJ 34. Using the Hess's law and the enthalpies of the given reactions, calculate the enthalpy of the following oxidation reaction between CuO and HCl: 2CuO(s) + 4HCl(g) → 2CuCl(s) + Cl2(g) + 2H2O(g), ΔH = ? 1) CuO(s) + H2(g) → Cu(s) + H2O(g), ΔH = -85 kJ 2) 2Cu(s) + Cl2(g) → 2CuCl(s), ΔH = -274 kJ 3) H2(g) + Cl2(g) → 2HCl(g), ΔH = -184 kJ So what is Hess's Law? This tutorial, you will be introduced to Hess's Law, as well as the equation that goes along with this concept. In addition, you will further master this concept by going through numerous example problems. Russian Chemist and Physicist Germain Hess developed the concepts of thermochemistry and physical chemistry. He introduced the concept known as Hess's Law of Constant Heat of Summation or Hess's Law for short. Hess's law has to do with net enthalpy in a reaction or set of thermochemical reactions and processes. Overall, it states that the total enthalpy change of a reaction is the sum of all the changes, no matter the number of steps or stages in the reaction (i.e. net enthalpy and the number of steps is independent of each other). There are some requirements that the reaction has to follow in order to use Hess's Law. For example, if there are multiple steps to the reaction, each equation must be balanced. Also, all the steps of the reaction must start and end at constant temperatures and pressures in order to keep reaction conditions constant. To put this definition into mathematical terms, here is the Hess's Law equation: $\Delta Hnet = \Sigma \Delta Hr$ net enthalpy change = $\Delta Hnet$ the sum of each enthalpy change steps = ΔHr Enthalpy change, ΔH, can be defined as the amount of heat absorbed or released during a reaction. Remember that enthalpy is only equal to the heat of reaction when the pressure and volume of the system are constant. In each individual step of a multistep reaction, there is a beginning and ending temperature, which is found by finding the difference between the final product enthalpy and the beginning reactant enthalpy ($\Delta Hnet$). That is Hess's Law equation. Now that we understand the concept and equation of Hess's Law, let's expand on our knowledge with practice problems. These word problems may ask for some manipulation of reactions (i.e. changing the direction of reactions, multiplication, division), but the general idea is the same for all Hess's Law problems. Let's go through some examples below! Find the net enthalpy change ($\Delta Hnet$) of the reaction below, given the reaction steps and their ΔH values. Overall reaction: N2H4(l) + H2(g) → 2NH3 (g)(i) N2H4(l) + CH4O(l) → CH2O(g) + N2(g) + 3H2(g)(ii) ΔH = −37kJ/mol(ii) N2(g) + 3H2(g) → 2NH3(g) ΔH = −46kJ/mol(iii) CH4O(l) → CH2O(g) + H2(g) ΔH = −65kJ/mol To make sure all the steps given are necessary for the overall reaction, add the equations and cross off repeated compounds to make a overall equation. However, if we do this step with the reactions as they are, we do not have N2H4 on the wrong side as well as extra compounds. Because of this, we can analyze if one, or more than one, of the steps, go in the opposite direction. Since reaction (i) is the only one with N2H4(l), which is a reactant in the overall equation, it is assumed that it is going in the correct direction. Next, reaction (ii) has the product 2NH3(g) on the right side, so that equation remains the same as well. In the above attempt to find the overall equation, the hydrogen gas from equations (i) and (ii) cancel each other out, meaning the hydrogen gas from reaction (iii) is the only one left to make it to the overall equation, which belongs on the left. Because of this, we can flip the hydrogen reactants and products to go the backward direction, because the reaction is going in the opposite way, the enthalpy also becomes the "opposite". If you change the direction of a reaction, the reciprocal of the enthalpy becomes the new enthalpy. The "new" equation steps look like this: (i) N2H4(l) + CH4O(l) → CH2O(g) + N2(g) + 3H2(g)(ii) ΔH = (-37 kJ/mol)(ii) N2(g) + 3H2(g) → 2NH3(g) ΔH = (-46 kJ/mol)(iii) CH2O(g) + H2(g) → CH4O(l) ΔH = +65kJ/mol With reaction (iii) switched the method of adding all the equations results in the correct overall equation: N2H4(l) + H2(g) → 2NH3 (g) Now that we have the official equation, we can use Hess's Law equation to solve. $\Delta Hnet = \Sigma \Delta Hr = $ (-37 kJ/mol) + (-46 kJ/mol) + 65 kJ/mol = −18kJ/mol Find the net enthalpy change ($\Delta Hnet$) of the reaction below, given the reaction steps and their ΔH values. Overall reaction: CS2(l) + 3O2(g) → CO2(g) + 2SO2(g) Now the same methods as above, we check if all the steps reactions are going in the correct direction. Reaction (i) has the desired CO2(g) product, which means it can remain unchanged. Reaction (iii) has CS2(l) as a product, but is a desired reactant in the overall reaction; therefore, we flip this reaction and use the reciprocal ΔH value. As for reaction (ii), the direction is correct because the 2SO2(g) are going in the correct direction; however, there is also an extra S(s) that needs to be canceled out. This can be fixed by multiplying reaction (ii) by a factor of 2. If you multiply(or divide) this, you also have to multiply (or divide) the ΔH value by the same coefficient. (i) C(s) + O2(g) → CO2(g) ΔH = -393.5 kJ/mol(ii) S(s) + O2(g) → SO2(g) ΔH = -296.8 kJ/mol(iii) C(s) + 2 S(s) → CS2(l) ΔH = +90. kJ/mol With reactions (ii) and (iii) manipulated, the method of adding all the equations results in the correct overall reaction: CS2(l) + 3O2(g) → CO2(g) + 2SO2(g) Now that we have the official equation, we can use Hess's Law equation to solve. $\Delta Hnet = \Sigma \Delta Hr = $ (-393.5 kJ/mol) + 2(-296.8 kJ/mol) + (-90 kJ/mol) = -1075 kJ/mol In the previous post, we talked about heat capacity, specific heat, and the formula for the relationship between these and the change in temperature. Below are some additional practice examples of heat capacity problems. When solving a problem related to heat capacity and heat transfer, remember that most of the time, it is assumed the heat is not lost, and it only flows from the object with a higher temperature to the colder one: The second expression adding the lost and gained heat is easier to use because you do not need to remember which one has a negative sign. 1. How much heat does it take to increase the temperature of a 540.6-g sample of Fe from 20.0 °C to 84.3 °C? The specific heat of iron = 0.450 J/g °C. 2. Calculate the specific heat capacity of a metal if a 17.0 g sample requires 481 J to change the temperature of the metal from 25.0 °C to 67.0 °C? 3. Calculate the energy of combustion for one mole if butane if burning a 0.367 g sample of butane (C4H10) has increased the temperature of a bomb calorimeter by 7.73 °C. The heat capacity of the bomb calorimeter is 2.36 kJ/ °C. 4. How many joules of energy is required to melt 40.0 g of ice at 0 °C? The heat of fusion (ΔHfus) for ice is 334.0 J/g. 5. How many kJ of energy does it take to change 36.0 g of ice at -15.0 °C to water at 0. °C? The specific heat of ice is 2.10 J/g°C and the heat of fusion (ΔHfus) for ice is 334.0 J/g. The terms endothermic and exothermic are used to describe the heat transfer that occurs during different processes including chemical reactions. In general, the focus in this chapter is going to be on heat transfer between molecules or molecular entities of a chemical reaction. Remember, heat is the transfer of thermal energy between two bodies that are at different temperatures. So, in a broader perspective, it may be used for any object. The system is essentially what we are interested in, while everything else is the surroundings. In chemistry, the system is the reaction and the changes in energy, heat, work or enthalpy are always looked at from the perspective of the system/reaction. For example, the combustion reaction of methane (CH4), the main component of natural gas, releases a large amount of heat, and therefore, it is an exothermic reaction (exo means "out", so heat flows out). 2H2O(l) + ΔHrxn = -891 kJ The ΔH is the enthalpy change, which, remember is equal to the heat (q) of the reaction under constant pressure. A chemical reaction that gives off heat to its surroundings, releases a negative ΔH, the enthalpy sum of the products is smaller than the one for products. Endothermic reactions (endo means "within," so heat flows in), on the other hand, absorb heat from surroundings and have a positive ΔH. For example, the formation of nitrogen monoxide (NO) from oxygen and nitrogen gases is an endothermic reaction that occurs during thunderstorms at high temperatures: N2(g) + O2(g) → 2NO(g) ΔH = +180.5 kJ Another example is the use of "instant cold" packs although these are not usually based on a chemical reaction, but rather on dissolving a salt such an ammonium nitrate (NH4NO3) or urea (CH4N2O). The salt is initially stored in a sealed plastic bag surrounded by water. When we break the container, the salt dissolved and dissociates in water breaking the ionic bonds which requires energy. This energy is obtained from the surroundings, and our body in this case. To summarize that: An endothermic reaction absorbs heat from the surroundings and has a positive ΔH Under constant pressure, the ΔH of a reaction is the amount of heat (q) absorbed or evolved in the reaction. Check Also Download the Testbook APP & Get Pass Pro Max FREE for 7 Days10,000+ Study NotesRealtime Doubt Support71000+ Mock TestsRankers Test Series+ more benefitsDownload App Now Hess's law states that the total enthalpy change of a reaction is independent of the pathway between the initial and final states. John M Land Photography Inc / Getty Images Hess's law states that the energy change in an overall chemical reaction is equal to the sum of the energy changes in the individual reactions comprising it. In other words, the enthalpy change of a chemical reaction (the heat of reaction) does not depend on the pathway between the initial and final states. The law is a variation of the first law of thermodynamics and conservation of energy. Because Hess's law holds true, it's possible to break a chemical reaction into multiple steps and use the standard enthalpies of formation to find the overall energy of a chemical reaction. Standard enthalpy tables are compiled from empirical data, usually around using calorimetry. Using these tables, it's possible to calculate whether or not a more complex reaction is thermodynamically favorable or not. In addition to calculating the enthalpy change of a reaction rather than directly measuring it, Hess's law is used to: Find electron affinities based on theoretical lattice energy. Calculate heat change of phase transitions. Calculate lattice energy when a substance changes allotropes. Find the heat of formation of an unstable intermediate in a reaction. Chakrabarty, D.K. (2001). An Introduction to Physical Chemistry. Mumbai: Alpha Science. pp. 34–37. ISBN 1-84265-059-9. Leicester, Henry M. (1951). "Germain Henri Hess and the Foundations of Thermochemistry". The Journal of Chemical Education. 28 (11): 581–583. doi:10.1021/ed028p581 Share — copy and redistribute the material in any medium or format for any purpose, even commercially. 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For example, other rights such as publicity, privacy, or moral rights may limit how you use the material. A calorimeter is a device for measuring heat flow. One application of calorimetry is measuring the heat of reaction. So, how does heat transfer between bodies? Therefore, if we plug in the expression $\Delta E = q + w$ in the enthalpy change equation, we get that the enthalpy change (ΔH) is equal to the heat when the pressure is constant. Once we have this settled, the rest of measuring the enthalpy of a reaction is similar to what we discussed about heat capacity and heat flow from one object to another. Remember, the key to solving problems in thermochemistry is the assumption that the heat is not lost, and it only flows from one object to the other: So, for calorimetry, we treat the reaction as one object or a system and the calorimeter as another object. The calorimeter can also be viewed as the surroundings. This does not change the principle that the heat flows between the reaction and the calorimeter. qrxn = qcal The heat transfer to/from the calorimeter is determined by its temperature change and the heat capacity: qcal = Ccal x ΔT The heat capacity of the calorimeter is determined experimentally and is already known when measuring the heat/enthalpy of a reaction. Although the calorimeters are designed for highly accurate measurement of heat transfers, a simple "coffee-cup" or "styrofoam-cup" calorimeter is often used in general chemistry laboratories. The styrofoam insulates well, and it serves the purpose of creating a closed system. So, the heat released or absorbed by the reaction changes the temperature of the water in the solution. This allows for calculating the heat of the solution, which is numerically equal to the heat of the reaction: qsolution = −qrxn qrxn solution is calculated as: qsoln = Cs m ΔT where Cs is the specific heat capacity of water (4.18 J/g oC), m is the mass of the solution, and ΔT is the temperature change of the solution. For example, When a 50.0 mL of 1.0 M HCl solution is mixed with 50.0 mL of 1.0 M NaOH solution in a coffee-cup calorimeter, the temperature of the resultant solution increases from 22.0 to 28.0 oC. Calculate the enthalpy change for the reaction in kJ/mol HCl assuming that the density of all solutions is 1 g/mL. We have an exothermic reaction because the temperature of the solution went up, and assuming the calorimeter does not lose any heat, we write: qrxn = − qsolution Next, calculate the qsoln: qsoln = Cs m ΔT where Cs is the specific heat capacity of water (4.18 J/g oC), the ΔT = 28.0 − 22.0 = 6.00 oC. The mass of the solution is the sum of the two solutions, and because each solution is 50.0 g, it is 100. g. Therefore, qsoln = 4.18 J/g oC (100. {(soln.)}}) × \mathOelta T_s = \left({100} \right) {\{{rm (soln)}}} ... the problem asks to determine the ΔH per one mole of HCl! So, we need to first calculate how many moles of HCl were in the 50.0 mL of 1.0 M HCl solution. For this, we use the molarity formula: \[{{\rm{n}}\]{{\rm{(HCl)}}} = {{\rm n} = } \], {{\rm{1}}}\frac{{{{\rm{mol}}} }}{\cancel{{{\rm{L}}}}} \times {{\rm{0}}}{{\rm{.0500}}} \, \cancel{{{\rm{L}}}}\] This corresponds to 2.51 kJ: \[{{\rm(2)}}\frac{{{\rm{51}}}}{\cancel{{{\rm(mol)}}}}\] corresponds to these specific amounts of reactants. However, the problem asks to determine the ΔH per one mole of HCl! So, we need to first calculate how many moles of HCl were in the 50.0 mL of 1.0 M HCl solution. For this, we see the molarity: \[{{\rm{n}}}\]{{\rm{(HCl)}}} = {{\rm n} = } \], {{\rm{1}}}\frac{{{{\rm{mol}}}}}{\cancel{{{\rm{L}}}}} \times {{\rm{0}}}{{\rm{.0500}}}\, \cancel{{{\rm{L}}}}\ = {{\rm{0}}}{{\rm{.0500}}}\, {{\rm{mol}}}\] This corresponds to 2.51 kJ, therefore, 1 mole of HCl would release: 2.51 kJ/0.0500 mol = 50.2 kJ/mol And this is the enthalpy change of the reaction per one mol of HCl: ΔH = −50.2 kJ/mol Hess's law of constant heat summation, also known simply as Hess's law, is a relationship in physical chemistry and thermodynamics.[1] Hess's law of constant heat summation, also known simply as Hess's law, is a relationship in physical chemistry and thermodynamics named after Germain Hess, a Swiss-born Russian chemist and physician who published it in 1840. The law states that the total enthalpy change during the complete course of a chemical reaction is independent of the sequence of steps taken.[2][3] Hess's law is now understood as an expression of the fact that the enthalpy of a chemical process is independent of the path taken from the initial to the final state (i.e. enthalpy is a state function). According to the first law of thermodynamics, the enthalpy change in a system due to a reaction at constant pressure is equal to the heat absorbed (or the negative of the heat released), which can be determined by calorimetry for many reactions. The values are usually stated for reactions with the same initial and final temperatures and pressures (while conditions are allowed to vary during the course of the reactions). Hess's law can be used to determine the overall energy required for a chemical reaction that can be divided into synthetic steps that are individually easier to characterize. This affords the compilation of standard enthalpies of formation, which may be used to predict the enthalpy change in complex synthesis. Hess's law states that the change of enthalpy in a chemical reaction is the same regardless of whether the reaction takes place in one step or several steps, provided the initial and final states of the reactants and products are the same. Enthalpy is an extensive property, meaning that its value is proportional to the system size.[4] Because of this, the enthalpy change is proportional to the number of moles participating in a given reaction. In other words, if a chemical value change takes place by several different routes, the overall enthalpy change is the same, regardless of the route by which the chemical change occurs (provided the initial and final condition are the same). If this were not true, then one could violate the first law of thermodynamics. Hess's law allows the enthalpy change (ΔH) for a reaction to be calculated even when it cannot be measured directly. This is accomplished by performing basic algebraic operations based on the chemical equation of reactions using previously determined values for the enthalpies of formation. Combination of chemical equations leads to a net or overall equation. If the enthalpy changes are known for all the equations in the sequence, their sum will be the enthalpy change for the net equation. If the net enthalpy change is negative (ΔH net < 0 {\displaystyle \Delta H_{\text{net}}}