

Thermochemistry flashcard



**ASSIGN
BUSTER**

The question asks us to determine ΔE , given information about q and w . Plan We first determine the signs of q and w (Table 5. 1) and then use Equation 5. , $\Delta E = q + w$, to calculate ΔE . Chemistry, The Central Science, 12th Edition Theodore L. Brown; H.

Eugene Lemma, Jar. ; Bruce E. Burette; Catherine J. Murphy; and Patrick Woodward © 2012 Pearson Education, Inc. 9 Sample Exercise 5.

2 Relating Heat and Work to Changes of Internal Energy Continued Solve

Heat is transferred from the system to the surroundings, and work is done on the system by the surroundings, so q is negative and w is positive: $q = -1150$ J and $w = 480$ kJ].

Thus, $\Delta H = q + w = (-1150 \text{ J}) + (480 \text{ J}) = -670 \text{ J}$ The negative value of tells us that a net quantity of 670 J of energy has been transferred from the system to the surroundings. Comment You can think of this change as a decrease of 670 J in the net value of the system's energy bank account (hence, the negative sign); 1150 J is withdrawn in the form of heat while 480 J is deposited in the form of work. Notice that as the volume of the gases decreases, work is being done on the system by the surroundings, resulting in a deposit of energy.

Practice Exercise Calculate the change in the internal energy for a process in which a system absorbs 140 J of heat from the surroundings and does 85 J of work on the surroundings.

Answer: +55 J Chemistry, The Central Science, 12th Edition Theodore L. Brown; H Eugene Lemma, Jar. ; Exchange of Heat between System and

Surroundings ; When heat is absorbed by the system from the surroundings, the process is endothermic. 10 endothermic. ; When heat is released by the system into the surroundings, the process is exothermic.

State Functions Usually we have no way of knowing the internal energy of a system; finding that value is simply too complex a problem. 11 ; However, we do know that the internal energy of a system is independent of the path by which the system achieved that state. – In the system depicted in Figure 5. 9, the water could have reached room temperature from either direction. ; Therefore, internal energy is a state function. ; It depends only on the present state of the system, not on the path by which the system arrived at that state.

; And so, ΔE depends only on Initial and Final. ; However, q and w are not state functions. Whether the battery is shorted out or is discharged by running the fan, its ΔE is the same. – But q and w are different in the two cases. Work Usually in an open container the only work done is by a gas pushing on the surroundings (or by the surroundings pushing on the gas). 13 We can measure the work done by the gas if the reaction is done in a vessel that has been fitted with a piston: $w = -P\Delta V$ Enthalpy ; If a process takes place at constant pressure (as the majority of processes we study do) and the only work done is this pressure-volume work, we can account for heat flow during the process by measuring the enthalpy of the system.

Enthalpy is the internal energy plus the product of pressure and volume: $H = E + PV$ 2012 Pearson Education, Inc. Thermosetting 14 ; When the system changes at constant pressure, the change in enthalpy, ΔH , is $\Delta H = \Delta E + P\Delta V$;

This can be written $\Delta H = \Delta H + p\Delta V$ Since $\Delta E = q + w$ and $w = -P\Delta V$, we can substitute these into the enthalpy expression: $\Delta H = \Delta E + P\Delta V$ $\Delta H = (q + w) - w$ $\Delta H = q$; So, at constant pressure, the change in enthalpy is the heat gained or lost. Thermochemistry 2012 Pearson 15 Endothermic and Exothermic; A process is endothermic when ΔH is positive. A process is endothermic when ΔH is positive.; A process is exothermic when ΔH is negative.

16 Sample Exercise 5.3 Determining the Sign of ΔH Indicate the sign of the enthalpy change, ΔH , in these processes carried out under atmospheric pressure and indicate whether each process is endothermic or exothermic: (a) An ice cube melts; (b) 1.00 mol of butane (C_4H_{10}) is combusted in sufficient oxygen to give complete combustion to CO_2 and H_2O . Analyze Our goal is to determine whether ΔH is positive or negative for each process.

Because each process occurs at constant pressure, the enthalpy change equals the quantity of heat absorbed or released, $\Delta H = q_p$. Plan We must predict whether heat is absorbed or released by the system in each process.

Processes in which heat is absorbed are endothermic and have a positive sign for ΔH ; those in which heat is released are exothermic and have a negative sign for ΔH . Solve In (a) the water that takes up the ice cube is the system. The ice cube absorbs heat from the surroundings as it melts, so ΔH is positive and the process is endothermic.

In (b) the system is the 1.00 mol of butane and the oxygen required to combust it. The combustion of butane in oxygen gives off heat, so ΔH is negative and the process is exothermic.

Molten gold poured into a mold solidifies at atmospheric pressure. With the gold defined as the system, is the solidification an exothermic or endothermic process? Answer: In order to solidify, the gold must cool to below its melting temperature. It cools by transferring heat to its surroundings. The air around the sample would feel hot because heat is transferred to it from the molten gold, meaning the process is exothermic.

You may notice that solidification of a liquid is the reverse of the melting we analyzed in the exercise. As we will see, reversing the direction of a process changes the sign of the heat transferred. Enthalpy of Reaction The change in enthalpy, ΔH , is the enthalpy of the products minus the enthalpy of the reactants: $\Delta H = H_{\text{products}} - H_{\text{reactants}}$ Thermosetting 2012 Pearson Education, 17 This quantity, ΔH , is called the enthalpy of reaction, or the heat of reaction. The Truth about Enthalpy . Enthalpy is an extensive property. 2.

ΔH for a reaction in the forward direction is equal in size, but opposite in sign, to ΔH for the reverse reaction. 3. ΔH for a reaction depends on the state of the products and the state of the reactants. Thermosetting 2012 Pearson Education, Inc.

18 Sample Exercise 5. 4 Relating ΔH to Quantities of Reactants and Products Continued Hydrogen peroxide can decompose to water and oxygen by the reaction $2 \text{H}_2\text{O}_2(\text{l}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$ $\Delta H = -196 \text{ kJ}$ Calculate the quantity of heat released when 5.00 g of $\text{H}_2\text{O}_2(\text{l})$ decomposes at constant pressure. Answer: -14.

4 kJ [Bruce E. Burette; Catherine].

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Murphy; and Patrick Ward Calorimeter Since we cannot know the exact enthalpy of the reactants and products, we measure ΔH through calorimeter, the measurement of heat flow. Thermochemistry 2012 19 Heat Capacity and Specific Heat The amount of energy required to raise the temperature of a substance by 1 K (1 °C) is its heat capacity. We define specific heat capacity (or simply specific heat) as the amount of energy required to raise the temperature of 1 g of substance by 1 K (or 1 °C).

20 Specific heat, then, is
$$\text{Specific heat} = \frac{\text{heat transferred}}{\text{mass} \times \text{temperature change}}$$
 $q = m \times c \times \Delta T$

Constant Pressure Calorimeter By carrying out a reaction in aqueous solution in a simple calorimeter such as this one, one can indirectly measure the heat change for the system by measuring the heat change for the water in the calorimeter. Thermochemistry 2012 Pearson Because the specific heat for water is well known (4.184 J/g-K), we can measure ΔH for the reaction with this equation: $q = m \times c \times \Delta T$ Sample Exercise 5.6 Measuring ΔH Using a Coffee-Cup Calorimeter When a student mixes 50 ml of 1.0 M HCl and 50 ml of 1.0 M NaOH in a coffee-cup calorimeter, the temperature of the resultant solution increases from 21.0 °C to 27.5 °C. Calculate the enthalpy change for the reaction in kJ/mol HCl, assuming that the calorimeter loses only a negligible quantity of heat, that the total volume of the solution is 100 ml, that its density is 1.0 g/ml, and that its specific heat is 4.18 J/g-K.

0 M NaOH in a coffee-cup calorimeter, the temperature of the resultant solution increases from 21.0 °C to 27.5 °C. Calculate the enthalpy change for the reaction in kJ/mol HCl, assuming that the calorimeter loses only a negligible quantity of heat, that the total volume of the solution is 100 ml, that its density is 1.0 g/ml, and that its specific heat is 4.18 J/g-K.

DC to 27.5 °C. Calculate the enthalpy change for the reaction in kJ/mol HCl, assuming that the calorimeter loses only a negligible quantity of heat, that the total volume of the solution is 100 ml, that its density is 1.0 g/ml, and that its specific heat is 4.18 J/g-K.

18 J/g-K. Analyze Mixing solutions of HCl and NaOH results in an acid-base reaction: $\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{NaCl(aq)}$ We need to calculate the heat produced per mole of HCl, given the temperature increase of the solution, the number of moles of HCl and NaOH involved, and the density and specific heat of the solution.

Plan The total heat released can be calculated using Equation 5. 23. The number of moles of HCl consumed in the reaction must be calculated from the volume and molarity of this substance, and this amount is then used to determine the heat produced per mol HCl.

Solve Because the total volume of the solution is 100 ml, its mass is The temperature change is Using Equation 5. 23, we have $(100 \text{ ml})(1.0 \text{ g/ml}) = 100 \text{ g}$ $\Delta T = 27.5 \text{ }^\circ\text{C} - 21.0 \text{ }^\circ\text{C} = 6.5$

$5 \text{ }^\circ\text{C} = 6.5 \text{ K}$ $q_{\text{rxn}} = -C_{\text{sp}} m \Delta T = -(4.18 \text{ J/g}\cdot\text{K}) = -2.70 \times 10^3 \text{ J} = -2.7 \text{ kJ}$

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Eugene Lemma, Jr. ; Bruce E.

Burette; Catherine J. Murphy; and Patrick Woodward O Sample Exercise 5.

6 Measuring ΔH_{rxn} Using a Coffee-Cup Calorimeter Continued Because the process occurs at constant pressure, To express the enthalpy change on a molar basis, we use the fact that the number of moles of HCl is given by the product of the volume ($\text{ml} = 0.050 \text{ L}$) and concentration.