ANSWER KEY.

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Reaction Energy Virtual Lab

Chemistry and AP Chemistry | ANSWER KEY

Open activities.zornscience.org/reaction-energy-lab/ and follow each step. Record values as the lab shows them. Chemistry students: the Chemistry version says PE diagram and kJ, and its heating curve uses per-gram constants, so some last digits may differ.

Objective

Use the Reaction Energy Lab to read and build energy diagrams, classify processes as exothermic or endothermic, calculate the heat in each part of a heating curve, use ΔH in stoichiometry, and find ΔH from bond enthalpies, enthalpies of formation, and Hess's law.

NGSS alignment

HS-PS1-4 (energy change in a reaction depends on bond energy), HS-PS3-4 (energy flow toward thermal equilibrium). Practices: developing and using models, using mathematics and computational thinking. Crosscutting concept: energy and matter.

AP alignment

Topics 6.1, 6.2, and 6.5 to 6.9.

Meet the simulation

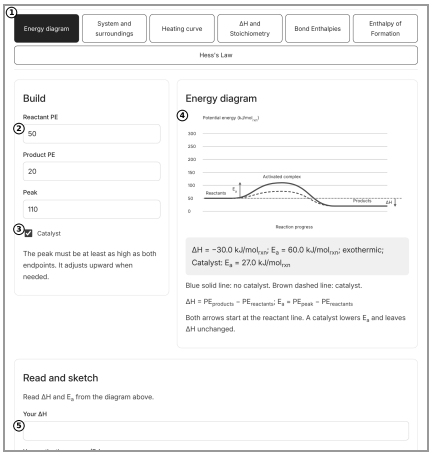


Figure 1. The Energy diagram mode. (1) Mode tabs: Energy diagram, System and surroundings, Heating curve, ΔH and Stoichiometry, Bond Enthalpies, Enthalpy of Formation, Hess's Law. (2) Type the reactant, product, and peak energies. (3) The Catalyst box adds a lower dashed path. (4) The diagram, with E_a and ΔH arrows that both start at the reactant line. (5) Read and sketch: type your answers and press Check.

Pre-lab

1. In an exothermic reaction, does the system gain or lose energy? What is the sign of ΔH, and what happens to the temperature of the surroundings?

The system loses energy to the surroundings. ΔH < 0, and the surroundings get warmer.

2. In Hess's law, what happens to ΔH when you reverse a reaction? When you multiply it by 2?

Reversing changes the sign of ΔH. Multiplying by 2 doubles ΔH.

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Part A. Energy diagram

3. Choose Energy diagram. Enter each set of energies, then read ΔH and E_a from the diagram and check them in Read and sketch.

Reactant PE	Product PE	Peak	ΔH (kJ/mol$_{rxn}$)	E_a (kJ/mol$_{rxn}$)	Exo or endo?
120	180	250	**+60.0**	**130**	**endothermic**
50	20	110	**−30.0**	**60.0**	**exothermic**

$\Delta H = PE_{products} - PE_{reactants}$; $E_a = PE_{peak} - PE_{reactants}$.

4. Keep the second set (50, 20, 110) and check Catalyst. Record the new E_a. What changes and what stays the same?

Catalyst E_a = 27.0 kJ/mol$_{rxn}$ (down from 60.0). ΔH stays −30.0 kJ/mol$_{rxn}$. A catalyst lowers the peak, not the reactant or product energies.

5. Sketch challenge: reactants = 25, ΔH = +50, E_a = 100. What product energy and peak energy should you enter? Check them in the lab.

Products = 25 + 50 = 75. Peak = 25 + 100 = 125.

Part B. System and surroundings

6. Choose System and surroundings. For each process, predict exothermic or endothermic, then press Start process and record what the surroundings temperature does.

Process	My prediction	Surroundings temperature
Dissolving KOH	**exothermic**	**rises (21.0 to 34.5 °C)**
Instant cold pack: NH_4NO_3 dissolving	**endothermic**	**falls**
Water vapor condensing on a cold glass	**exothermic**	**rises**
$Ba(OH)_2 \cdot 8H_2O + 2NH_4Cl$: frost forms	**endothermic**	**falls**

When the surroundings warm up, the system released energy (exothermic). When they cool down, the system absorbed energy (endothermic).

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Part C. Heating curve

7. Choose Heating curve. Heat 15.0 g of water from −20 °C to 120 °C. Fill in each segment, then record the total q.

Segment	What happens	Setup	q
A	**ice warms**	**15.0 g × 2.06 × 20.0**	**618 J**
B	**melting**	**(15.0 ÷ 18.02) mol × 6.01 kJ/mol**	**5.00 kJ**
C	**liquid warms**	**15.0 g × 4.18 × 100**	**6.27 kJ**
D	**boiling**	**(15.0 ÷ 18.02) mol × 40.7 kJ/mol**	**33.9 kJ**
E	**steam warms**	**15.0 g × 2.01 × 20.0**	**603 J**

Total q = 4.64×10^4 J = 46.4 kJ. Boiling (D) takes by far the most energy.

8. Use Read the curve. What are the melting and boiling points? In which segments are two phases present, and does the kinetic energy or the potential energy change there?

Melting point 0 °C, boiling point 100 °C. Segments B and D have two phases. The temperature is constant, so KE stays the same and the energy goes into PE (separating particles).

Part D. ΔH and Stoichiometry

9. Choose $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, $\Delta H = -890.\ kJ/mol_{rxn}$. Find q for burning 8.00 g of CH_4 and check it.

8.00 g ÷ 16.04 g/mol = 0.4988 mol = 0.4988 mol_{rxn}
q = 0.4988 mol_{rxn} × (−890. kJ/mol_{rxn}) = −444 kJ

10. Choose $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, $\Delta H = -572\ kJ/mol_{rxn}$. Find q for 10.0 g of H_2. What answer do you get if you forget the coefficient 2?

10.0 ÷ 2.016 = 4.960 mol H_2 ÷ 2 = 2.480 mol_{rxn}
q = 2.480 × (−572) = -1.42×10^3 kJ.
Forgetting the 2 gives -2.84×10^3 kJ, twice too much.

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Part E. Bond Enthalpies

11. Choose Bond Enthalpies. For each reaction, tally the bonds broken and formed, press Check, and record the sums and ΔH.

Reaction	Σ bonds broken (kJ)	Σ bonds formed (kJ)	ΔH (kJ/mol_{rxn})
$H_2 + Cl_2 \rightarrow 2HCl$	**436 + 242 = 678**	**2(431) = 862**	**−184**
$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O(g)$	**4(413) + 2(495) = 2642**	**2(799) + 4(463) = 3450**	**−808**
$N_2 + 3H_2 \rightarrow 2NH_3$	**941 + 3(436) = 2249**	**6(391) = 2346**	**−97**

ΔH = Σ(broken) − Σ(formed). The CH_4 estimate (−808) differs from the measured −802 because bond enthalpies are averages.

Part F. Enthalpy of Formation

12. Choose Enthalpy of Formation. Find $\Delta H°_{rxn}$ for $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ and for $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, using the lab's $\Delta H°_f$ table.

$CaCO_3$: [−635.1 + (−393.5)] − (−1206.9) = +178.3 kJ/mol_{rxn}
CH_4: [−393.5 + 2(−285.8)] − [−74.8 + 0] = −890.3 kJ/mol_{rxn}

Part G. Hess's Law

13. Choose the set $C(s) + ½O_2(g) \rightarrow CO(g)$.

Reverse or multiply the steps until the net reaction matches the target. What did you change, and what is ΔH?

Keep step 1 (−393.5) and reverse step 2 (+283.0). ΔH = −393.5 + 283.0 = −110.5 kJ/mol_{rxn}

14. Choose the set $2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$.

What did you change, and what is ΔH?

Multiply step 1 by 2: 2(−296.8) = −593.6. Keep step 2 (−197.8).
ΔH = −593.6 + (−197.8) = −791.4 kJ/mol_{rxn}

Analysis and conclusion

15. Claim, evidence, reasoning: why can bond enthalpies, enthalpies of formation, and Hess's law all be used to find the same kind of quantity, ΔH?

Claim: ΔH is a state function, so any path from the same reactants to the same products gives the same ΔH. Evidence: the CH_4 combustion gave −890.3 from formation data, close to −890. used in Part D, and Hess steps added to −110.5 for CO. Reasoning: each method breaks the reaction into steps (breaking bonds, forming compounds from elements, or known reactions) whose energies add up.

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