

King of Dumbs

Level: ♠ Technical

Why doesn't diamond turn into graphite?

Breaking bonds costs energy and making them pays it back, but almost no reaction starts without climbing a hill. Why diamond lasts and water works as a battery.

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<https://kingofdumbs.com/en/lessons/chemical-reactions-energy>

What you'll learn

- Why some reactions give off heat and others take it in: the energy in bonds
- Activation energy: the hill to climb before reacting
- Stable isn't the same as fast: why diamond lasts
- Oxidation and reduction: splitting water and burning it is charging and draining a battery

Before this lesson

If something sounds like gibberish, these lessons explain it from scratch:

1. [What is a wall made of? Atoms and elements without fear](#)
2. [Joules, watts or calories? Energy and power without fear](#)

Symbols you'll meet

Δ change: how much it moved (♠) · E_a subscript: the label underneath (♦♠) · H° ° for standard, not degrees (♠) · Σ sum: add them all up (♠) · $\approx$ approximately equal (♠) · $|x|$ absolute value: drop the sign (♠) · · dot: multiply (or dot product) (♠) · $\rightarrow$ arrow: turns into / tends to (♦♠) · J/(mol·K) compound units (♦♠) · e the number e (≈ 2.718) (♠) · log logarithm: how many times do I multiply? (♠) · ν nu: coefficient or neutrino (♠)

A diamond ring can pass from grandmother to granddaughter for centuries and stay the same. But a chemist would tell you that diamond “should” turn into graphite, the stuff in a pencil, because graphite is more stable. And it doesn’t.

This lesson’s dumb question is: **why doesn’t diamond turn into graphite?** The answer is the same one that explains why the gas on your stove doesn’t catch fire by itself, why food keeps longer in the fridge and why hydrogen stores energy like a battery. Every reaction has a bill, the energy it gives off or asks for, and a hill to climb before it starts.

Pick your card: ♣ if you’re starting from zero, ♦ if a simple formula doesn’t scare you, ♠ if you want the notation. You can switch at any time.

Breaking costs, joining pays

♦ → ♠ From intermediate to technical

In ♦ we said	In ♠ we write
bond energy	$D(X-Y)$: bond dissociation enthalpy in the gas phase, in kJ/mol
the heat of the reaction	reaction enthalpy $\Delta_r H$ (heat at constant pressure); H is a state function
486 kJ is left over	$\Delta_r H \approx \Sigma D(\text{broken}) - \Sigma D(\text{formed}) = 1370 - 1856 = -486 \text{ kJ}$
exothermic	$\Delta_r H < 0$
endothermic	$\Delta_r H > 0$
averages	$\bar{D}$, averaged over many molecules: errors of tens of kJ/mol
the energy of forming each substance from its elements	standard enthalpy of formation $\Delta_f H^\circ$ (0 for an element in its most stable form)
you can do the sum without bonds	Hess's law: $\Delta_r H^\circ = \Sigma \nu \Delta_f H^\circ(\text{products}) - \Sigma \nu \Delta_f H^\circ(\text{reactants})$, with ν the coefficients

$$\Delta_r H \approx \Sigma D(\text{broken}) - \Sigma D(\text{formed})$$

$$2H_2 + O_2: 1370 - 1856 = -486 \text{ kJ}$$

$$\Delta_r H^\circ = \Sigma \nu \Delta_f H^\circ(\text{prod}) - \Sigma \nu \Delta_f H^\circ(\text{react})$$

$$D(H-H) = 2 \cdot \Delta_f H^\circ(H) = 436 \text{ kJ/mol}$$

$$D(O=O) = 498.3; \quad \bar{D}(O-H) = 463.5$$

The reaction enthalpy is estimated with $\Delta_r H \approx \Sigma D(\text{broken}) - \Sigma D(\text{formed})$ (OpenStax, *Chemistry 2e*, 7.5). The D values in the table are averages, and the estimate is usually off by tens of kJ: in OpenStax's example, $CO + 2 H_2 \rightarrow CH_3OH(g)$ comes out at -107 kJ from bonds and -90.5 kJ from enthalpies of formation.

For water the estimate works because $H-H$ and $O=O$ aren't averages: each is one specific molecule. With the NIST-JANAF tables at 298.15 K, $D(H-H) = 2 \Delta_f H^\circ(H, g) = 2 \times 217.999 = 436.0 \text{ kJ/mol}$, $D(O=O) = 2 \Delta_f H^\circ(O, g) = 498.3 \text{ kJ/mol}$, and the average $O-H$ in water is $[2 \Delta_f H^\circ(H) + \Delta_f H^\circ(O) - \Delta_f H^\circ$

$(\text{H}_2\text{O}, \text{g})] / 2 = 463.5 \text{ kJ/mol}$.

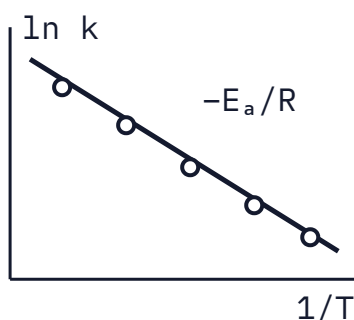
With those D values, the estimate gives exactly $\Delta_f H^\circ(\text{H}_2\text{O}, \text{g}) = -241.826 \text{ kJ/mol}$: it's the same sum by another route. Hess's law guarantees it: since H is a state function, $\Delta_r H^\circ = \sum \Delta_f H^\circ(\text{products}) - \sum \Delta_f H^\circ(\text{reactants})$ doesn't depend on the intermediate steps (OpenStax, 5.3).



The hill: why almost nothing starts by itself

♦ → ♠ From intermediate to technical

In ♦ we said	In ♠ we write
activation energy	E_a , in kJ/mol
energy profile	energy against the reaction coordinate; the top is the transition state (‡)
The share of collisions that reaches the top	$e^{(-E_a/RT)}$, with $R = 8.314 \text{ J/(mol}\cdot\text{K)}$ and T in kelvin
runs 4.6 times slower	rate constant $k = A \cdot e^{(-E_a/RT)}$ (Arrhenius), with A the frequency factor
rises enormously with temperature	$\ln k = \ln A - (E_a/R)(1/T)$: a straight line with slope $-E_a/R$
every 10 degrees, twice as fast	$k_2/k_1 = \exp[(E_a/R)(1/T_1 - 1/T_2)] = 1.92$ with $E_a = 50 \text{ kJ/mol}$, from 298 to 308 K
catalyst	another mechanism with $E_a' < E_a$ and the same $\Delta_r H$



$$k = A \cdot e^{(-E_a/RT)}$$

$$\ln k = \ln A - (E_a/R) \cdot (1/T)$$

$$k_2/k_1 = \exp[(E_a/R)(1/T_1 - 1/T_2)]$$

$$E_a = 50 \text{ kJ/mol}, 298 \rightarrow 308 \text{ K}: \times 1.92$$

The Arrhenius equation is $k = A \cdot e^{(-E_a/RT)}$ (OpenStax, 12.5). The factor $e^{(-E_a/RT)}$ measures what share of collisions arrives with enough energy; A, the frequency factor, is related to how often molecules collide and how they are oriented. Taking logs, $\ln k = \ln A - (E_a/R)(1/T)$: measure k at several temperatures, plot $\ln k$ against $1/T$ and you get a straight line whose slope gives E_a . OpenStax does it with $2 \text{ HI} \rightarrow \text{H}_2 + \text{I}_2$ and gets $E_a = 180 \text{ kJ/mol}$.

Between two temperatures, $k_2/k_1 = \exp[(E_a/R)(1/T_1 - 1/T_2)]$. With $E_a = 50 \text{ kJ/mol}$, from 298.15 to 308.15 K it gives 1.92; with 100 kJ/mol, 3.70 (our own calculation, assuming A stays constant). The rule of doubling every 10 °C isn't a law: it depends on E_a .

A catalyst offers another mechanism, with a lower transition state. Since H is a state function, the start and finish don't move: in OpenStax's diagrams with and without a catalyst, «both curves begin and end at the same energies» (12.7).

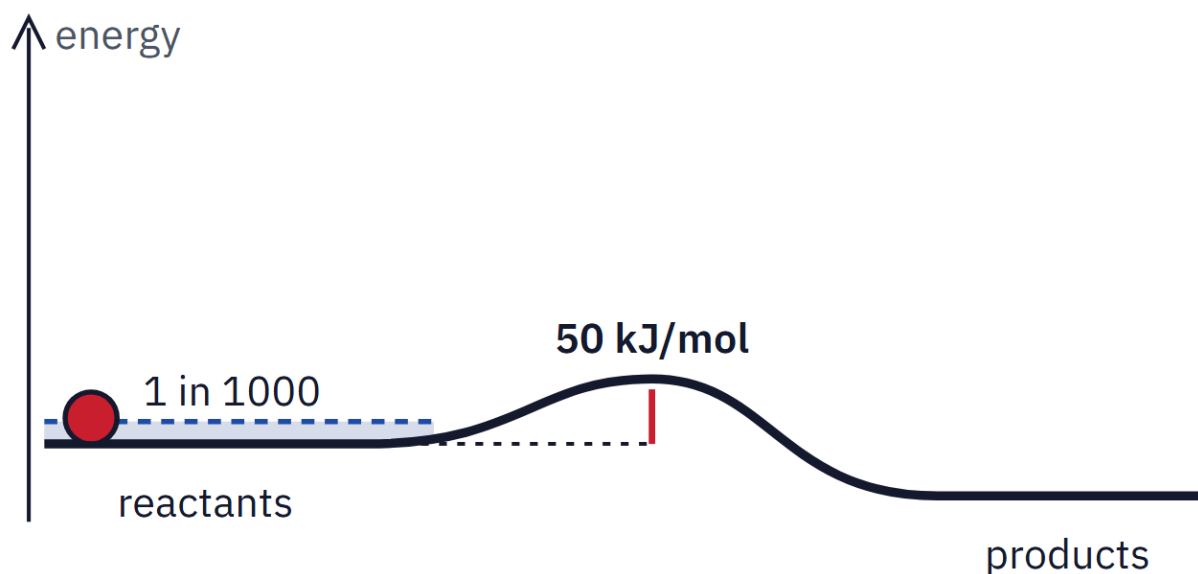
The King of Dumbs says:

It's not enough for a reaction to be worth it: first you have to climb the hill. That's why paper doesn't burst into flames on the shelf.



Try it: climb the hill

Move the height of the hill and the temperature. The blue band marks how high 1 in 1000 collisions gets: when it's over the top, the reaction runs fast. Try 190 kJ/mol, close to the hill measured in diamond powder, and turn up the temperature: even at 1,500 °C the band doesn't get there, but look how much faster it goes.



Height of the hill 50 kJ/mol

Temperature 25 °C (room)

Collisions that reach the top **1 in 5.7×10^8**

Compared with 25 °C the same

If it took a minute at 25 °C, it would take 1 min

A 50 kJ/mol hill at 25 °C (room): 1 in 5.7×10^8 collisions reach the top.



Interactive widget. Try it on the web: <https://kingofdumbs.com/en/lessons/chemical-reactions-energy>

Diamond: stable isn't the same as fast

♦ → ♠ From intermediate to technical

In ♦ we said	In ♠ we write
Gibbs energy	$G = H - TS$; at constant T and p, $\Delta G = \Delta H - T\Delta S$
disorder	entropy S, in J/(K·mol)
If the Gibbs energy goes down, the reaction can go by itself	$\Delta_r G < 0$: spontaneous process
goes down by 2.9 kJ/mol	$\Delta_r G^\circ = -2.90$ kJ/mol (standard state, 298.15 K)
the heat it gives off	$\Delta_r H^\circ = -1.89$ kJ/mol: $ \Delta_r H^\circ < \Delta_r G^\circ $ because $\Delta_r S^\circ > 0$
a powder of tiny diamonds	ultradisperse diamond (UDD): nanodiamond
one face	{110} and {111} faces: planes of the crystal (Miller indices)
the boundary between diamond and graphite moves	interface velocity $v = A \cdot e^{(-E/RT)}$ (Butenko and others)
thermodynamically unstable but kinetically stable	$\Delta_r G^\circ < 0$, but $k \approx 0$ at 298 K
metastable	a local minimum of G separated from the global minimum by a barrier

$$\Delta_r G^\circ = \Delta_r H^\circ - T\Delta_r S^\circ$$

$$= -1.89 - 298.15 \cdot 0.00336 = -2.89 \text{ kJ/mol}$$

$$v = A \cdot e^{(-E/RT)}, A = 74 \text{ nm/s}, E = 188 \text{ kJ/mol}$$

$$v(1873 \text{ K}) \approx 1.5 \text{ nm/h}$$

$$v(298 \text{ K}) \approx 1 \text{ nm} / 4 \times 10^{23} \text{ years}$$

$$E_a\{110\} = 728 \text{ kJ/mol (Davies, Evans)}$$

From OpenStax's appendix G, diamond → graphite has $\Delta_r H^\circ = -1.89$ kJ/mol and $\Delta_r S^\circ = 5.740 - 2.38 = +3.36$ J/(K·mol). So $\Delta_r G^\circ = \Delta_r H^\circ - T\Delta_r S^\circ = -1.89 - 298.15 \times 0.00336 = -2.89$ kJ/mol, which matches the -2.90 in the table. Careful: the heat given off is $\Delta_r H$, -1.89; what decides the direction is $\Delta_r G$.

The kinetics were measured by Davies and Evans (*Proc. R. Soc. Lond. A* 328, 413, 1972) between 1850 and 2000 °C, at zero pressure: $E_a = 728 \pm 50$ kJ/mol on the 110 faces. For the 111 faces their abstract gives two figures that don't agree with each other: 253 kcal/mol and 1159 kJ/mol, when 253 kcal is 1059 kJ. They propose that the slow step is pulling a single atom off the surface: one held by two bonds on the 110 faces and by three on the 111.

Butenko and others (*J. Appl. Phys.* 88, 4380, 2000), with nanodiamond between 1370 and 1870 K, fit $v = A \cdot e^{(-E/RT)}$ with $E = 45 \pm 4$ kcal/mol (188 kJ/mol) and $A = 74$ nm/s. They say their formula works «in and around the indicated temperature range» and put the difference from Davies and Evans down to another mechanism at lower temperatures.

If you extrapolate their formula, $v(298\text{ K}) \approx 8 \times 10^{-32}$ nm/s: one nanometre in about 4×10^{23} years, between 5×10^{20} and 4×10^{26} with their ± 4 kcal/mol. With Davies and Evans's E_a , the exponential factor at 25 °C is 10^{110} times smaller than at 1900 °C. These are extrapolations far beyond what was measured: only the order of magnitude counts. At ambient pressure, diamond is a local minimum of G separated from graphite by a barrier; at very high pressures, diamond becomes the stable one (OpenStax, 16.1).

Oxidising and reducing: water as a battery

♦ → ♠ From intermediate to technical

In ♦ we said	In ♠ we write
oxidation	losing e^- : the oxidation number goes up
reduction	gaining e^- : the oxidation number goes down
electrolysis	electrolytic cell: a voltage $ E > E^\circ_{\text{cell}} $ is imposed to drive a reaction with $\Delta_r G > 0$
at least 1.23 volts	$E^\circ = \Delta_r G^\circ / (nF) = 1.229 \text{ V}$, with $n = 2 \text{ e}^-$ per H_2O and $F = 96,485 \text{ C/mol}$ (Faraday constant)
285.8 kJ per mole	$\Delta_f H^\circ(\text{H}_2\text{O}, \text{l}) = -285.83 \text{ kJ/mol}$; $\Delta_r H / (nF) = 1.481 \text{ V}$, the thermoneutral voltage
the rest as heat from the surroundings	$T\Delta_r S^\circ = \Delta_r H^\circ - \Delta_r G^\circ = 285.8 - 237.1 = 48.7 \text{ kJ/mol}$
fuel cell	galvanic cell: anode $2 \text{ H}_2 \rightarrow 4 \text{ H}^+ + 4 \text{ e}^-$; cathode $\text{O}_2 + 4 \text{ H}^+ + 4 \text{ e}^- \rightarrow 2 \text{ H}_2\text{O}$
efficiency	maximum electrical work $-\Delta_r G$; the rest, $T\Delta_r S$, comes out as heat

anode: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ (ox.)

cathode: $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ (red.)

$$E^\circ = \Delta_r G^\circ / (nF) = 237.1 \text{ kJ} / (2F) = 1.229 \text{ V}$$

$$\Delta_r H / (nF) = 1.481 \text{ V}; \quad T\Delta_r S = 48.7 \text{ kJ/mol}$$

$$-\Delta_r G / -\Delta_r H = 83 \% \quad (\text{fuel cell})$$

In electrolysis (OpenStax, 17.7), the anode oxidises water, $2 \text{ H}_2\text{O}(\text{l}) \rightarrow \text{O}_2 + 4 \text{ H}^+ + 4 \text{ e}^-$, and oxygen goes from -2 to 0 . The cathode reduces, $2 \text{ H}^+ + 2 \text{ e}^- \rightarrow \text{H}_2$, and hydrogen goes from $+1$ to 0 . $E^\circ_{\text{cell}} = -1.229 \text{ V}$: you have to impose more than that.

A check with $\Delta_r G^\circ = -nFE^\circ$: $\Delta_f G^\circ(\text{H}_2\text{O}, \text{l}) = -237.141 \text{ kJ/mol}$ (NIST-JANAF) and $n = 2$, so $E^\circ = 237,141 / (2 \times 96,485.332) = 1.2289 \text{ V}$. With $\Delta_r H$ instead of $\Delta_r G$ you get 1.481 V , the thermoneutral voltage: above it, the cell needs no heat from outside. The gap between the two voltages is the heat the cell can draw from its surroundings, $T\Delta_r S = 48.7 \text{ kJ/mol}$, which matches $\Delta_r S^\circ = 163.3 \text{ J/(K}\cdot\text{mol)}$ from appendix G.

The other way round, in a fuel cell, the maximum electrical work is $-\Delta_r G = 237.1 \text{ kJ/mol}$, 83 % of what burning the hydrogen would give (our own calculation). OpenStax gives about 1.2 V per cell and a real efficiency of 50–75 %. The energy that comes out never exceeds what it cost to split the water: it's a chemical battery, not a source.

The King of Dumbs says:

Hydrogen isn't an energy mine: it's a battery. What you get out by burning it, you put in before by splitting the water.



Final challenge

1. The gas on your stove and the oxygen in the air sit together and don't burn until you light the burner. Why?

- a. Because burning gas doesn't give off energy
- b. Because first there's a hill to climb: it needs a push, like a spark
- c. Because the gas has no energy until it's heated

2. Graphite is more stable than diamond. Why doesn't a diamond ring turn into graphite?

- a. Because diamond is actually the most stable
- b. Because the hill between them is huge and at room temperature almost no atom gets across
- c. Because diamond and graphite are different elements

3. You split water with electricity and then burn the hydrogen that comes out. How much energy do you get?

- a. More than you spent: hydrogen is a source of energy
- b. At most what you spent: the water works as a battery
- c. None, because water doesn't burn



In three sentences

1. A reaction breaks bonds, which costs energy, and makes others, which give it back: if there's some left over it comes out as heat (exothermic), and if there's some missing you have to put it in (endothermic).
2. Almost no reaction starts by itself, because first there's a hill to climb, the activation energy: heat and catalysts help get over it.
3. That's why diamond lasts even though graphite is more stable, and why splitting water and burning the hydrogen is charging and draining a battery, never gaining energy.



Sources

- Bond energy and estimating ΔH : OpenStax, *Chemistry 2e*, [7.5](#). Enthalpy, exothermic and endothermic, Hess's law and enthalpies of formation: [5.3](#) and [appendix G](#). NIST-JANAF tables at 298.15 K: [H](#), [O](#), [water vapour](#) and [liquid water](#).
- Temperature and speed (the fridge): [OpenStax, 12.2](#). Activation energy and Arrhenius: [12.5](#). Catalysts: [12.7](#). Gas constant and Faraday constant (CODATA 2022): NIST, [R](#) and [F](#).
- Spontaneity and Gibbs energy: OpenStax, [16.1](#) and [16.4](#). Metastable diamond: [UW–Madison, Chem 109, D26.1](#).
- Diamond kinetics: G. Davies and T. Evans, *Proc. R. Soc. Lond. A* 328, 413 (1972), [doi:10.1098/rspa.1972.0086](#); Yu. V. Butenko and others, *J. Appl. Phys.* 88, 4380 (2000), [doi:10.1063/1.1289791](#); O. V. Sedelnikova and others, *Beilstein J. Nanotechnol.* 16, 887 (2025), [doi:10.3762/bjnano.16.67](#).
- Oxidation and reduction: [OpenStax, 4.2](#). Electrolysis of water: [17.7](#). Fuel cells: [17.5](#).

How this lesson was made: researched and written with help from Claude from the sources above, checked and cross-checked against a second source on 16 October 2026 and reviewed again on 17 October 2026. Figures without their own link are calculations from them and from the atoms and energy lessons. The drawings are our own.

Answers

1. **b)** Because first there's a hill to climb: it needs a push, like a spark. That's it. First the push; after that the flame pushes itself.
2. **b)** Because the hill between them is huge and at room temperature almost no atom gets across. Right. Stable isn't the same as fast.
3. **b)** At most what you spent: the water works as a battery. Exactly. Climbing the slope costs what going down gives back.

A King of Dumbs lesson, kingofdumbs.com. Offline version; the original, with widgets you can play with, is at: <https://kingofdumbs.com/en/lessons/chemical-reactions-energy>

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