

Energy, Heat and Work

What are energy, heat and work, and how are they related?

●○○ Beginner

Thermochemistry & Thermodynamics

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Before you start: it helps to understand [SI Units and Unit Conversions](#) and [Temperature Scales: Celsius, Fahrenheit and Kelvin](#).

What is it?

Energy is the capacity to do work or to transfer heat. It comes in two broad kinds:

- **Kinetic energy**, the energy of motion: $E_k = \frac{1}{2}mv^2$. The random motion of particles is **thermal energy**.
- **Potential energy**, energy stored because of position or arrangement: a raised object ($E_p = mgh$), charges near each other, or the **chemical energy** stored in the arrangement of atoms and electrons in bonds.

Energy is measured in **joules** (J). One joule is the energy of a force of one newton acting over one metre: $1 \text{ J} = 1 \text{ kg m}^2 \text{ s}^{-2}$. Other units you will meet:

Unit	In joules
kilojoule, kJ	10^3 J
calorie, cal	4.184 J (exactly)
food Calorie, Cal (= 1 kcal)	4184 J
kilowatt-hour, kWh	$3.6 \times 10^6 \text{ J}$
litre-atmosphere, L·atm	101.325 J

Energy moves between things in only two ways: as **heat** (q), a transfer caused by a **temperature difference**, or as **work** (w), a transfer caused by a **force acting through a distance**, such as a gas pushing back a piston.

KEY IDEA

Energy is **never created or destroyed**, only transferred or changed from one form to another. This is the **first law of thermodynamics**. For a chemical system:

$$\Delta U = q + w$$

The change in the system's **internal energy**, ΔU , equals the heat added to it plus the work done on it.

Why does it matter?

- **Every reaction involves energy.** Burning fuel, charging a battery and digesting food all convert chemical energy into heat, work or electrical energy.
- **It is the foundation of thermochemistry.** Enthalpy, Hess's law and Gibbs free energy are all built on the first law.
- **Everyday units.** Food labels give energy in kJ and kcal; electricity bills in kWh. Converting between them is the same dimensional analysis you already know.

How does it work?

1. System and surroundings

The **system** is the part we study (for example, the chemicals in a flask); the **surroundings** are everything else. Systems can be:

- **open:** matter and energy can cross the boundary (an open pan of boiling water);
- **closed:** energy can cross but matter cannot (a sealed flask);
- **isolated:** neither can cross (an ideal sealed, insulated flask).

2. Heat is not temperature

Temperature measures how hot something is: it reflects the **average** kinetic energy of its particles. **Heat** is **energy in transit** from a hotter object to a colder one. A bath of warm water has a lower temperature than a cup of boiling water, but it holds far more thermal energy, because it contains far more particles. Heat always flows from **higher to lower temperature**, until both are at the same temperature (thermal equilibrium).

3. Work done by a gas

When a gas expands against a constant external pressure, it pushes the surroundings back and does work on them:

$$w = -P_{\text{ext}} \Delta V$$

The minus sign means that when the gas **expands** (ΔV positive) the system **loses** energy (w negative). When the gas is **compressed**, w is positive. If the volume does not change, no

expansion work is done ($w = 0$).

4. Sign convention

Chemistry looks at everything **from the system's point of view**:

	Positive (+)	Negative (-)
q	Heat flows into the system	Heat flows out
w	Work is done on the system	Work is done by the system
ΔU	Internal energy increases	Internal energy decreases

∞ THINK OF IT LIKE THIS

Think of the system's internal energy as a bank account. Heat and work are the only two kinds of transaction: deposits are positive, withdrawals are negative. The balance (U) only changes by the sum of the transactions, and money is never printed or destroyed.

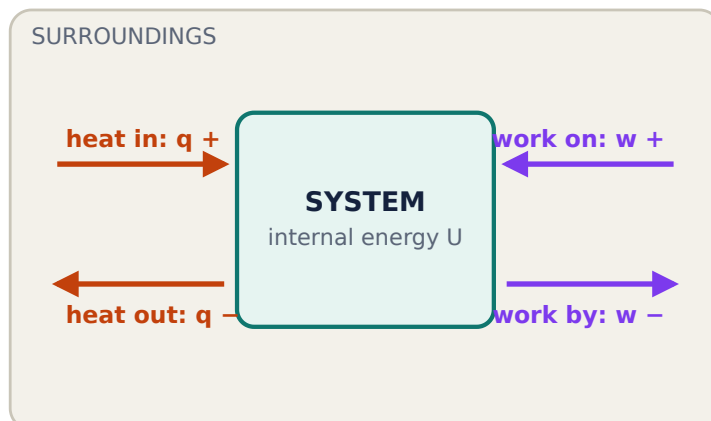
Q MORE PRECISELY

Internal energy is a **state function**: ΔU depends only on the start and end states, not on the route. Heat and work are **not** state functions: the same ΔU can be reached with different splits of q and w . Many physics books write the first law as $\Delta U = Q - W$, where W is work done **by** the system; it is the same law with a different sign convention. At constant pressure, the heat transferred equals the **enthalpy change**, $q_P = \Delta H$, which is why chemists usually measure ΔH (see [Enthalpy and Calorimetry](#)).

Visualise it

System and surroundings

Arrows into the system are positive.



$$\Delta U = q + w$$

Energy crosses the boundary of a system only as heat or work. Arrows into the system are positive.

Worked example

WORKED EXAMPLE: Kinetic energy

Question: A 0.145 kg baseball travels at 40.0 m/s. What is its kinetic energy?

$$\begin{aligned} E_k &= \frac{1}{2}mv^2 \\ &= \frac{1}{2} \times 0.145 \text{ kg} \times (40.0 \text{ m/s})^2 \\ &= 116 \text{ kg m}^2 \text{ s}^{-2} \\ &= 116 \text{ J} \end{aligned}$$

WORKED EXAMPLE: Converting energy units

Question: A snack bar provides 250 kcal. Express this in kilojoules.

$$\begin{aligned} 250 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} \\ = 1046 \text{ kJ} = 1.05 \times 10^3 \text{ kJ} \end{aligned}$$

 WORKED EXAMPLE: Work done by an expanding gas

Question: A gas expands from 2.00 L to 5.00 L against a constant external pressure of 1.50 atm. Calculate the work in joules.

Step 1 Change in volume: $\Delta V = 5.00 \text{ L} - 2.00 \text{ L} = 3.00 \text{ L}$

Step 2 Work:

$$\begin{aligned}w &= -P_{\text{ext}} \Delta V \\&= -1.50 \text{ atm} \times 3.00 \text{ L} \\&= -4.50 \text{ L atm} \\&\quad \times \frac{101.325 \text{ J}}{1 \text{ L atm}} \\&= -456 \text{ J}\end{aligned}$$

Step 3 The sign is negative: the gas does work **on the surroundings** and loses 456 J.

 WORKED EXAMPLE: The first law

Question: While expanding as in the last example, the gas absorbs 1.20 kJ of heat. What is ΔU ?

Step 1 Heat flows in, so $q = +1.20 \text{ kJ} = +1200 \text{ J}$. From the last example, $w = -456 \text{ J}$.

Step 2 First law:

$$\begin{aligned}\Delta U &= q + w \\&= (+1200 \text{ J}) + (-456 \text{ J}) \\&= +744 \text{ J}\end{aligned}$$

Step 3 The internal energy of the gas **increases** by 744 J: it gained more energy as heat than it lost as work.

Common mistake

 COMMON MISTAKE: Treating heat and temperature as the same thing

Temperature is a property of an object (in °C or K); heat is energy transferred (in J). A spark at 1000 °C carries very little heat; a warm bath at 40 °C can deliver a lot.

 COMMON MISTAKE: Getting the sign of work wrong

An **expanding** gas does work on the surroundings, so w is **negative**. Check: the system is spending energy, so its share must go down.

 COMMON MISTAKE: Leaving work in L·atm

$P\Delta V$ in atm $\times$ L gives L·atm, not joules. Multiply by 101.325 J/(L·atm) before adding w to a heat in joules.

 NOTATION NOTE

- U is internal energy (some books use E); q is heat; w is work.
- The symbol Δ (“change in”) always means final minus initial: $\Delta V = V_{\text{final}} - V_{\text{initial}}$.
- A food “Calorie” with a capital C is a kilocalorie: 1 Cal = 1 kcal = 4.184 kJ.

Remember this

 REMEMBER THIS

- Energy: kinetic ($\frac{1}{2}mv^2$) or potential (mgh , chemical); unit J; 1 cal = 4.184 J; 1 L·atm = 101.325 J.
- Energy crosses a boundary only as heat (q , driven by a temperature difference) or work (w , a force through a distance).
- Temperature is not heat: temperature measures average particle energy; heat is energy in transit.
- Gas expansion work: $w = -P_{\text{ext}}\Delta V$; no volume change means no expansion work.
- First law: $\Delta U = q + w$; positive means energy into the system.

References

1. Bureau International des Poids et Mesures (BIPM). The International System of Units (SI), 9th ed.; BIPM, 2019.
[Link](#)
2. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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