

AP CHEMISTRY

Unit 5 Thermochemistry

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Kinetic Energy and Potential Energy

- Kinetic energy - the energy of motion:

$$E_k = \frac{1}{2}mv^2$$

- Potential energy - the energy an object possesses by virtue of its position.
- Potential energy can be converted into kinetic energy, and vice versa. Example: a bicyclist riding up and down a hill.

Units of Energy

- SI Unit for energy is the joule, J:

$$\begin{aligned} E_k &= \frac{1}{2} m v^2 = \frac{1}{2} (2 \text{ kg})(1 \text{ m/s})^2 \\ &= 1 \text{ kg m}^2 \text{ s}^{-2} = 1 \text{ J} \end{aligned}$$

We sometimes use the calorie instead of the joule:

$$1 \text{ cal} = 4.184 \text{ J (exactly)}$$

A nutritional Calorie:

$$1 \text{ Cal} = 1000 \text{ cal} = 1 \text{ kcal}$$

Examples – Convert the following values:

- 1. 395 Calories to joules.**
- 2. 880. joules to calories.**
- 3. 307984 joules to Calories**

Examples – Convert the following values:

1. 395 Calories to joules.

1650000 J

2. 880. joules to calories.

210. cal

3. 307984 joules to Calories

73.6099 Cal

Systems and Surroundings

- System: part of the universe we are interested in.
- Surroundings: the rest of the universe.

- Force is a push or pull on an object.
- Work is the product of force applied to an object over a distance:
$$w = F \leftrightarrow d$$
- Energy is the work done to move an object against a force.
- Heat is the transfer of kinetic energy between two objects (depends on temperature and amount of substance).
- **Energy** is the capacity to do work or transfer heat.

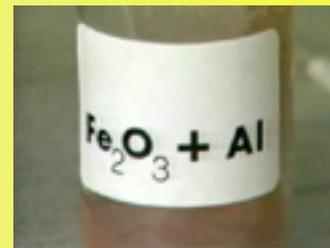
Internal Energy

- Internal Energy: total energy of a system.
- Cannot measure absolute internal energy.
- Change in internal energy,

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$

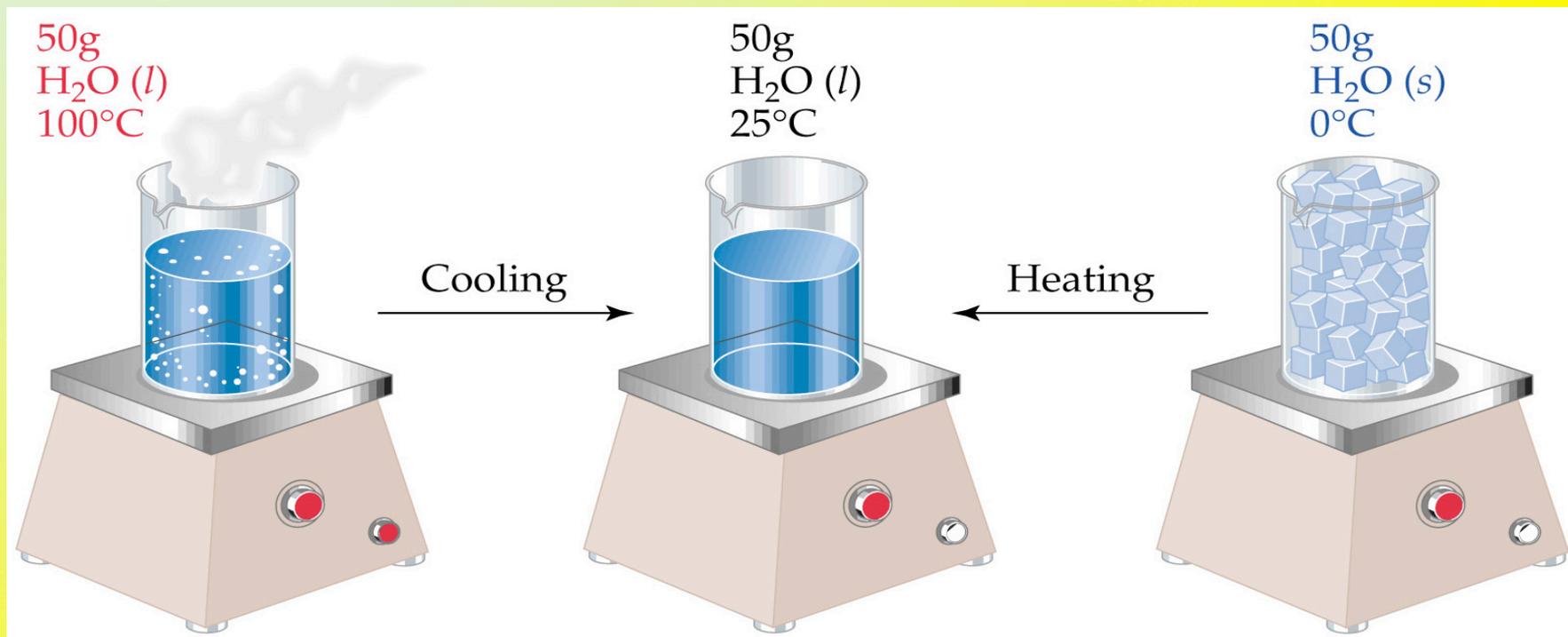
Exothermic and Endothermic Processes

- **Endothermic**: absorbs heat from the surroundings.
- **Exothermic**: transfers heat to the surroundings.
- An endothermic reaction feels cold.
- An exothermic reaction feels hot.



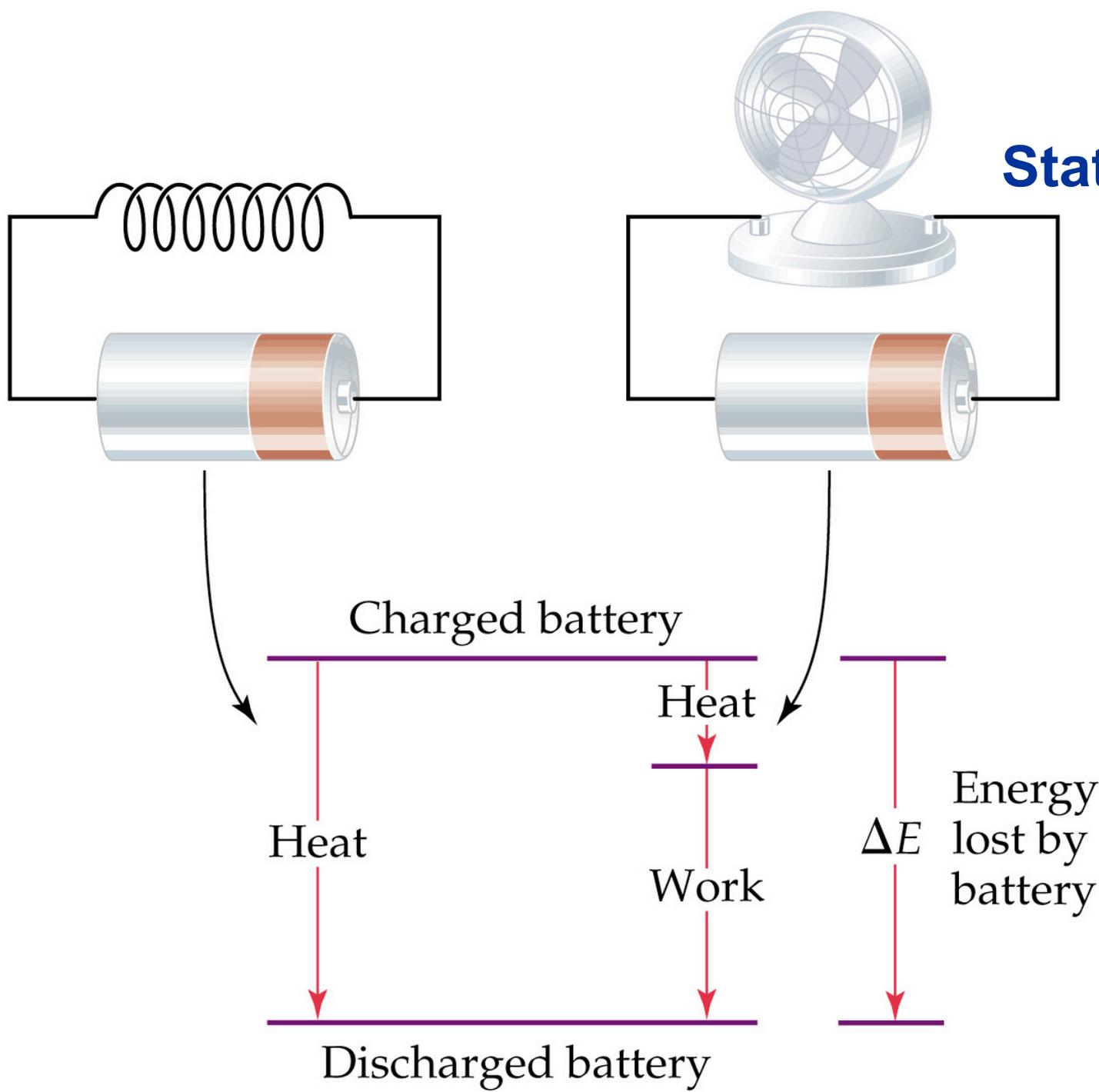
State Functions

- State function: depends only on the initial and final states of system, not on how the internal energy is used.



- Path function: depends on how the system gets from one set of circumstances to another

State Functions



- Chemical reactions can absorb or release heat.
- However, they also have the ability to do work.
- For example, when a gas is produced, then the gas produced can be used to push a piston, thus doing work.

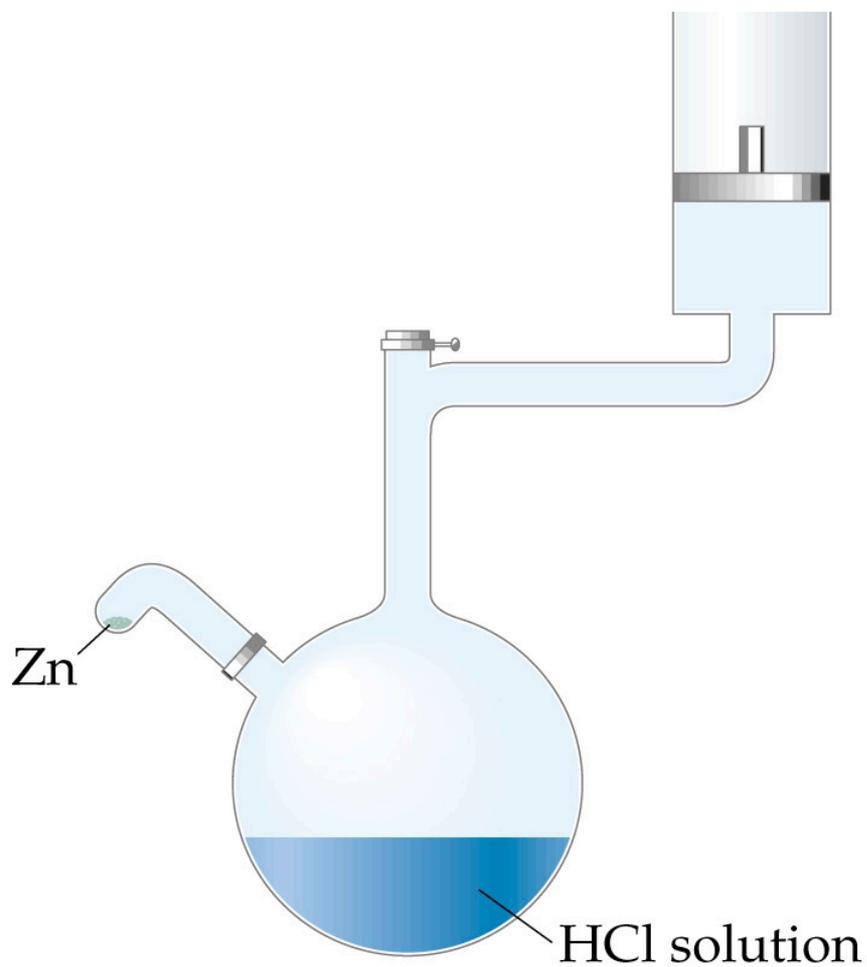


- The work performed by the above reaction is called *pressure-volume* work.
- When the pressure is constant,

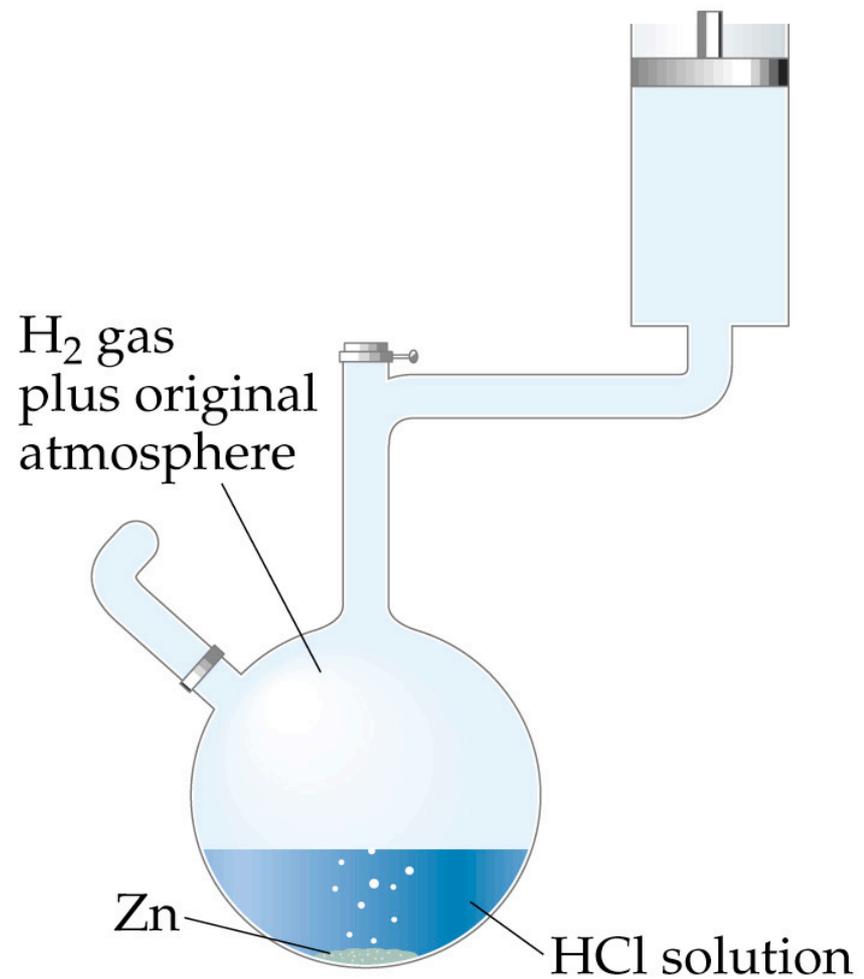
$$w = -P\Delta V$$



- 1 Joule = 1 L kPa



(a)



(b)

- Enthalpy, H : Heat transferred between the system and surroundings carried out under constant pressure.
- Enthalpy is a state function.
- When ΔH is positive, the system gains heat from the surroundings.
- When ΔH is negative, the surroundings gain heat from the system.

Surroundings

System



$$\Delta H > 0$$

Endothermic

Surroundings

System



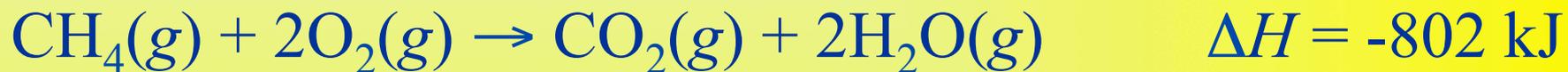
$$\Delta H < 0$$

Exothermic

- For a reaction:

$$\begin{aligned}\Delta H &= H_{\text{final}} - H_{\text{initial}} \\ &= H_{\text{products}} - H_{\text{reactants}}\end{aligned}$$

- Enthalpy is an *extensive* property (magnitude of ΔH is directly proportional to amount):



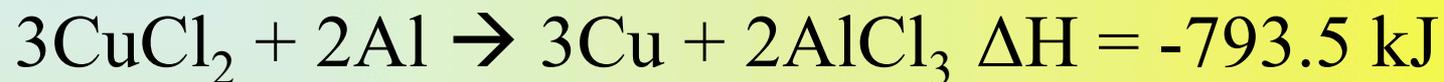
- When we reverse a reaction, we change the sign of ΔH :



- Change in enthalpy depends on state:



Examples



1. What is ΔH for the following reaction?

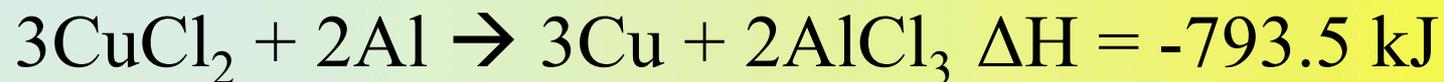


2. What is ΔH for the following reaction?

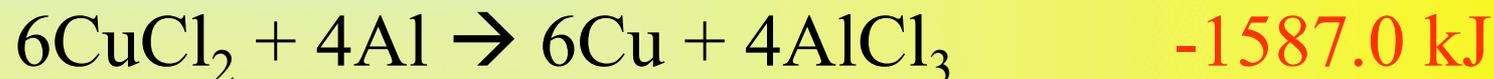


3. Given the equation at the top of the page, how much energy will be given off by the reaction of 2.07 g of Al with excess CuCl_2 ?

Examples



1. What is ΔH for the following reaction?



2. What is ΔH for the following reaction?



3. Given the equation at the top of the page, how much energy will be given off by the reaction of 2.07 g of Al with excess CuCl_2 ? **30.4 kJ**

Heat Capacity and Specific Heat

- Calorimetry = measurement of heat flow.
- Calorimeter = apparatus that measures heat flow.
- Heat capacity = the amount of energy required to raise the temperature of an object (by one degree). ($\text{J } ^\circ\text{C}^{-1}$)
- Molar heat capacity = heat capacity of 1 mol of a substance. ($\text{J mol}^{-1} \text{ } ^\circ\text{C}^{-1}$)
- Specific heat = specific heat capacity = heat capacity of 1 g of a substance. ($\text{J g}^{-1} \text{ } ^\circ\text{C}^{-1}$)

$$q = (\text{specific heat}) \times (\text{grams of substance}) \times \Delta T$$

- $q = mC\Delta T$

Constant Pressure Calorimetry

- Atmospheric pressure is constant!

$$mC\Delta T = -mC\Delta T$$

Object 1	Object 2
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$$\Delta H = q_P$$

$$q_{\text{rxn}} = -q_{\text{soln}}$$

$$= -(\text{specific heat of solution}) \times (\text{grams of solution}) \times \Delta T$$

- HEAT LOST = HEAT GAINED

Constant Pressure Calorimetry



Example

An unknown metal is to be analyzed for specific heat.

A 12.84-g piece of the metal is placed in boiling water at 101.5°C until it attained that temperature.

The metal is then removed from the boiling water and placed into a styrofoam cup containing 54.92 g of water at 23.2°C . The final temperature of the water was 26.1°C . What is the specific heat of the metal sample, in $\text{J g}^{-1} \text{ }^{\circ}\text{C}^{-1}$?

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$$.69 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1}$$

Example

25.0 mL of 0.100 M NaOH is mixed with 25.0 mL of 0.100 M HNO₃, and the temperature of the mixture increases from 23.1°C to 49.8°C. Calculate ΔH for this reaction in kJ mol⁻¹. Assume that the specific heat of each solution is equal to the specific heat of water, 4.184 J g⁻¹ °C⁻¹.

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-2230 kJ mol⁻¹