Auburn Mountainview

Karl Steffin, 2006

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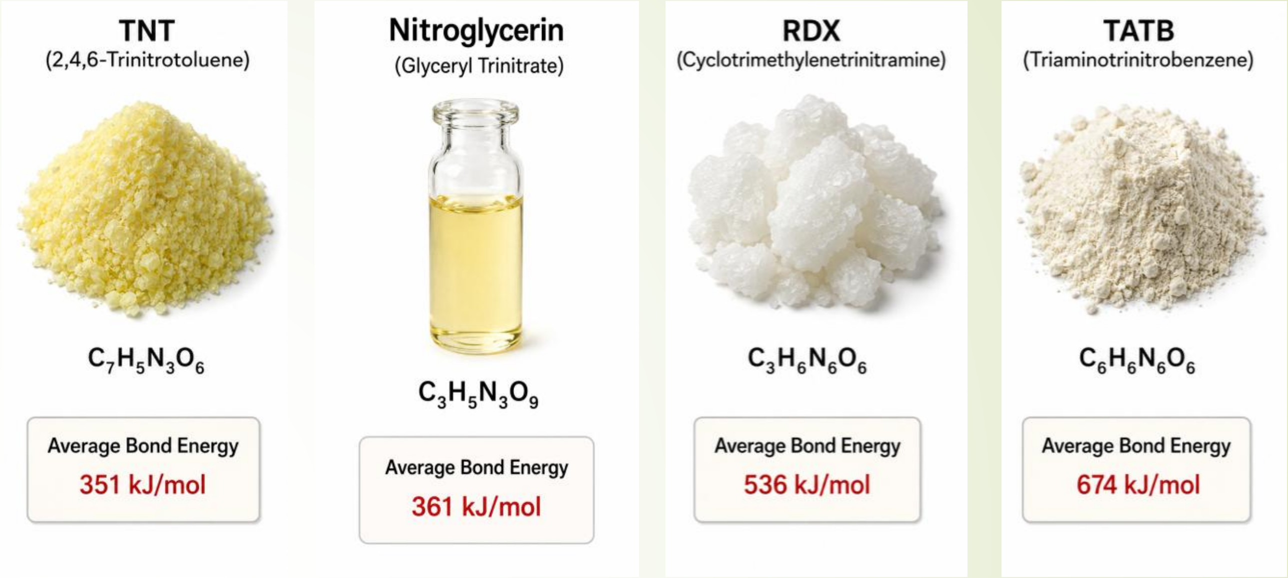
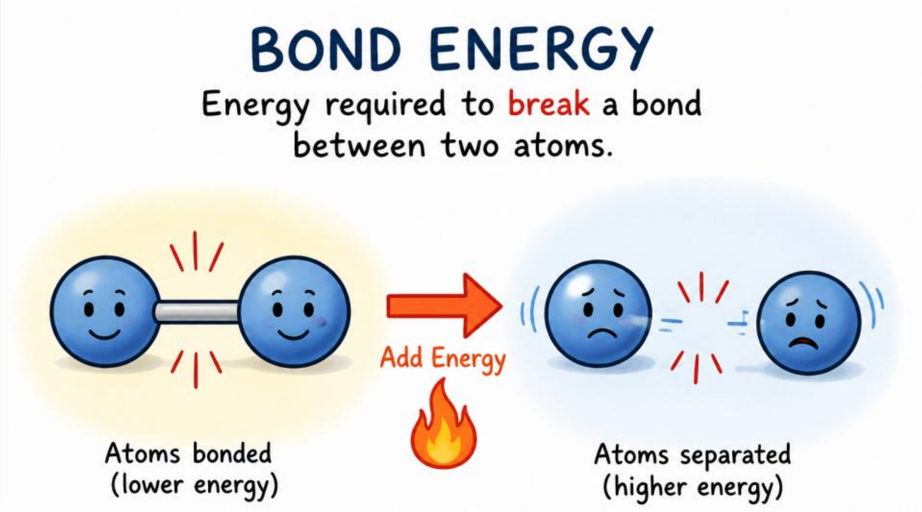


Thermochemistry

By end of this unit I can…

- A.1 Create and/or evaluate a claim about whether a reaction is endothermic or exothermic from experimental observations.
- A.2 Explain the relationship between the measured change in temperature of a solution and the energy transferred by a chemical reaction.
- A.3 Calculate energy changes in chemical reactions from calorimetry data.
- B.1 Create and/or evaluate a claim about the energy transferred as a result of a chemical reaction based on bond energies.

Bond Energy

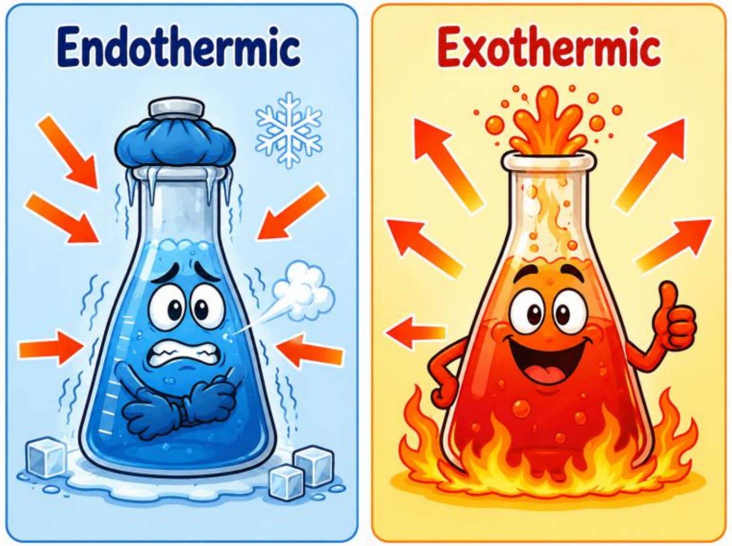


- Ionic Bonds disassociate in water.
  - Ions can then recombine to form new compounds.
- What about covalent reactions? Take a look at the following:

$$2H_2 + O_2 \rightarrow 2H_2O$$

- The H_2 and O_2 bonds were broken to form H_2O bonds.
- Bond Energy is the amount of energy needed break apart one mole of a specific bond in a substance's gas phase.
  - Measured in kJ per mole

Systems Versus Surroundings

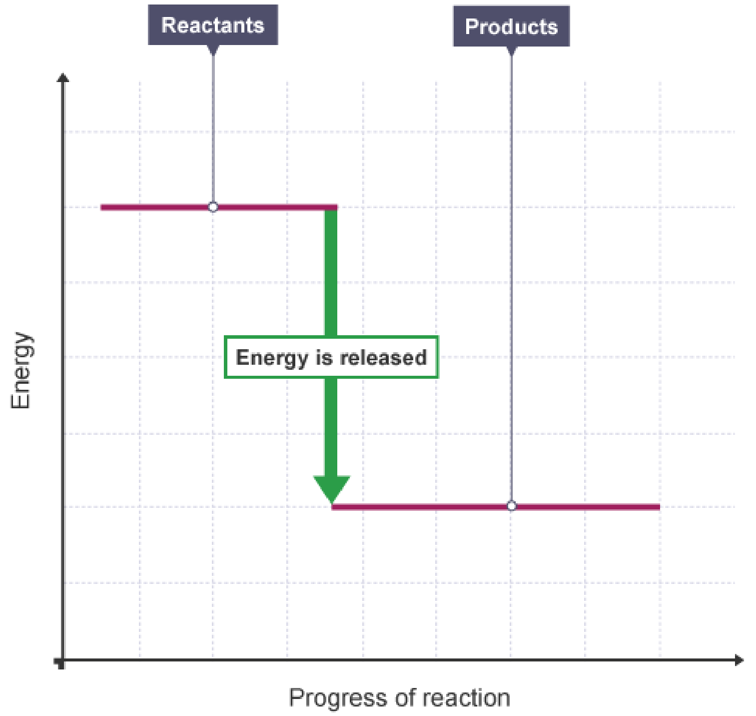


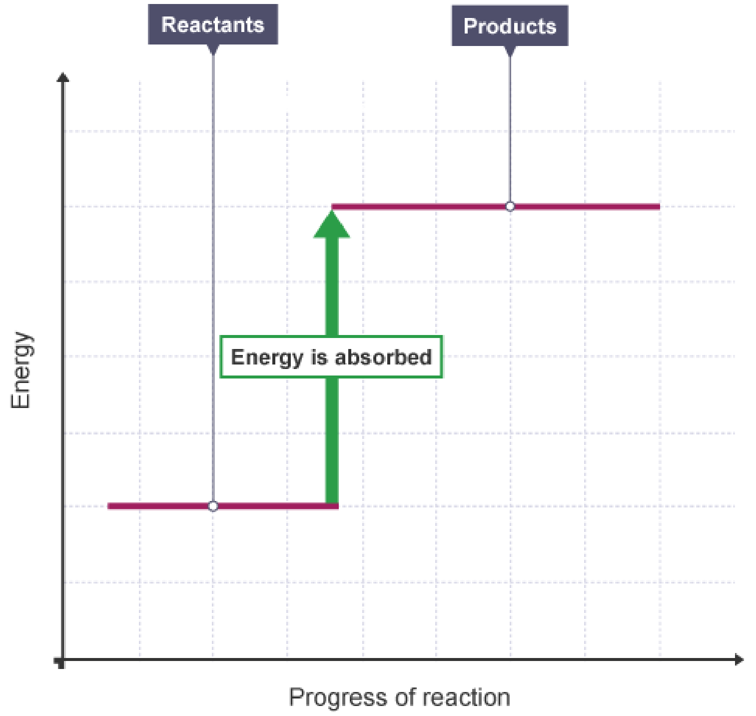
$$2H_2 + O_2 \rightarrow 2H_2O$$

- $2H_2$= 872-kJ/mol O_2= 498-kJ/mol $2H_2O$= 1852-kJ/mol
  - Net Change: 1852-1370= 482-kJ/mol
- This reaction has an excess of energy left over.
  - The bonds of the product(s) requires less energy than the reactant(s). Exothermic (out).
- If the product releases less energy the reaction is Endothermic (into).

Energy Diagrams

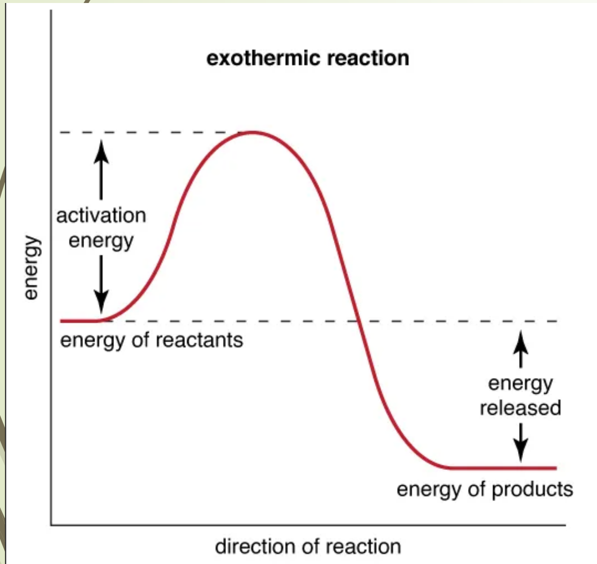
- The top graph shows that breaking the reactants bonds is a higher energy than the resulting products bonds: Energy is left over.
  - Exothermic Reaction
- The bottom graph shows that breaking the reactants bonds is a lower energy than the resulting products bonds: More energy is required to complete the reaction.
  - Endothermic Reaction

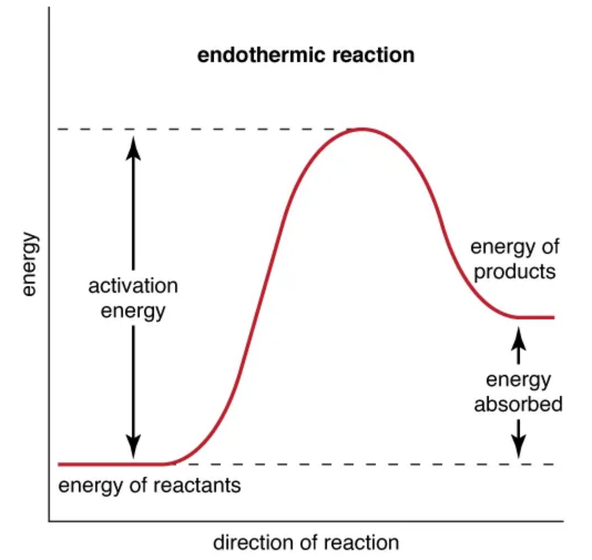


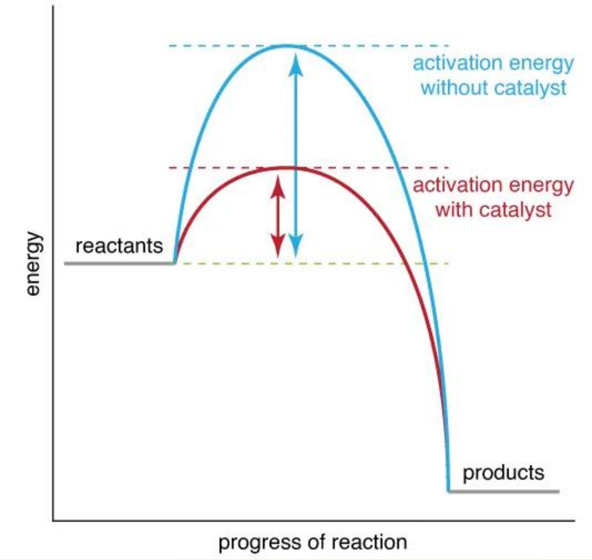


Activation Energy and Catalysts

- For a reaction to occur some energy is needed.
  - Heat: Increase collisions and break bonds.
  - Single Replacement: Reactive metals have lower activation energies.
  - Catalysts lower activation without being consumed.

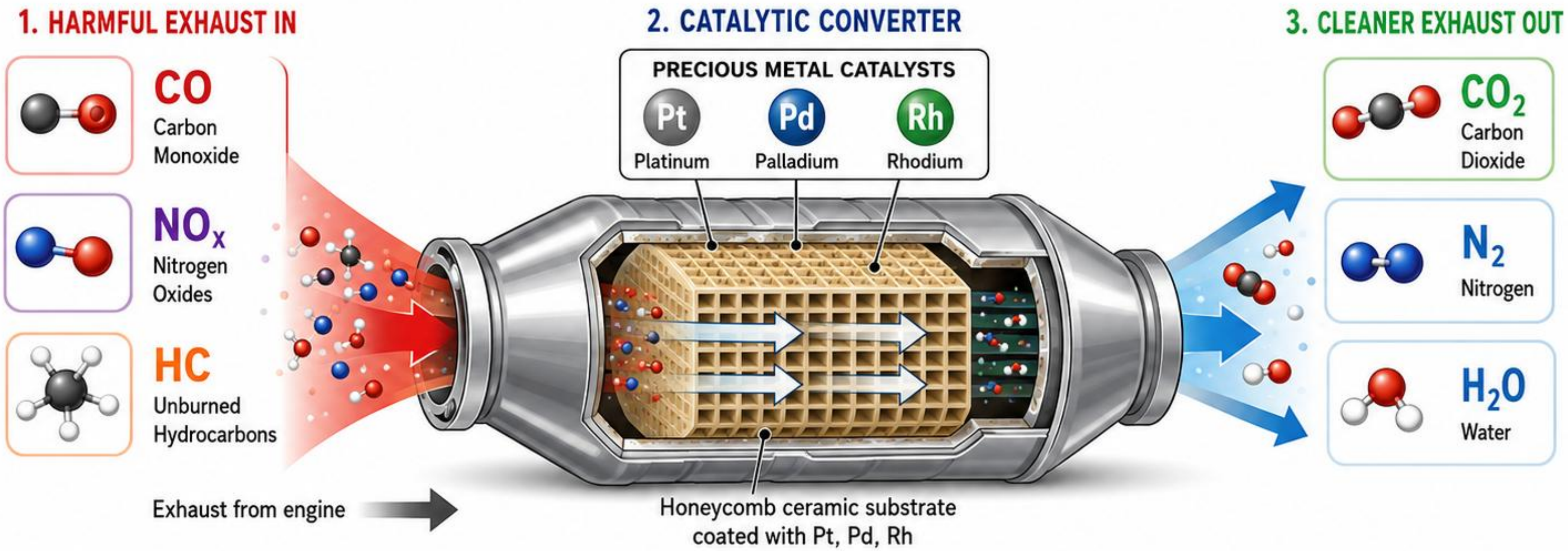






HOW A CATALYTIC CONVERTER WORKS

A catalytic converter uses precious metal catalysts to speed up chemical reactions that transform harmful exhaust gases into less harmful ones.



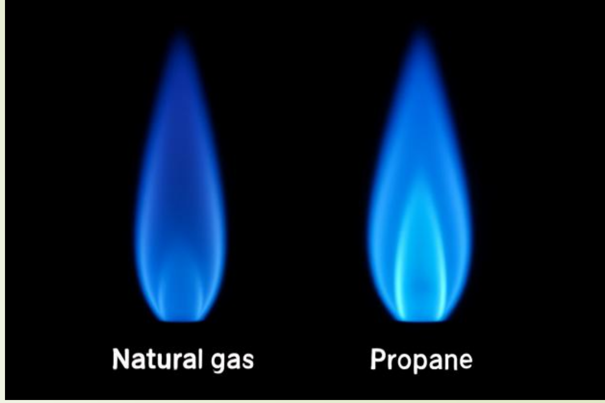
CHEMICAL REACTIONS (Simplified)

Carbon Monoxide Oxidation	$2CO + O_2 \rightarrow 2CO_2$	Carbon monoxide reacts with oxygen to form carbon dioxide.
Nitrogen Oxide Reduction	$2NO \rightarrow N_2 + O_2$	Nitrogen oxides break down to nitrogen and oxygen.
Hydrocarbon Oxidation	$C_xH_y + O_2 \rightarrow CO_2 + H_2O$	Unburned hydrocarbons react with oxygen to form carbon dioxide and water.

LEGEND

- Carbon (C)
- Oxygen (O)
- Nitrogen (N)
- Hydrogen (H)
- Honeycomb Ceramic Substrate

Enthalpy ($\Delta H°$)



- Enthalpy considers pressure and volume of the reaction.
  - (1-atm, 298.15-K)
  - Also shows a return to the starting point.
- The combustion of Propane:

$$C_3H_{8(g)} + 5O_2 \rightarrow 3CO_{2(g)} + 4H_2O_{(g)} + 2043\text{-kJ}$$

- To get the conditions back to starting point:

$$C_3H_{8(g)} + 5O_2 \rightarrow 3CO_{2(g)} + 4H_2O_{(g)} \qquad \Delta H° = -2043\text{-kJ}$$

Enthalpy Example

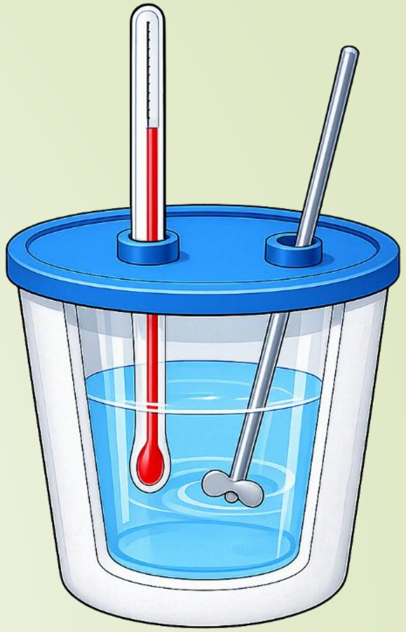
- How much heat will be released if 1.0-g of Hydrogen Peroxide decomposes? $2H_2O_2 \rightarrow 2H_2O + O_2$ $\Delta H° = -190 kJ/mol$

1.0-g

$$1-g\ H_2O_2\ x\ \frac{1-mol}{34-g}\ x\ \frac{190-kJ}{2-mol\ H_2O_2} =$$

2.8-kJ

Calorimetry



- Calorimetry is the study of heat flow and measurement.
  - $q = cm\Delta T$
- Done in an isolated environment (no heat loss).
  - Heat from reaction equals opposite the heat from system.

$$0 = q_{rxn} + q_{sys} \qquad \text{or} \qquad q_{rxn} = -q_{sys}$$

- Water is used to reliably measure change in temperature.
  - $c = 4.184\text{-J/g}^\circ\text{C}$

Calorimetry Example I (Partial)

- **5.75-g** of K_2CO_3 dissolves in 80.0-g of water. The temperature drops from 21.0-°C to 16.9-°C. Calculate ΔH°.
  - The systems change in temperature is given.

Solve for the Q of the water (system).

$$Q_{Water} = Cm\Delta T$$

$$Q_W = 80 - g \cdot 4.184 - \frac{J}{g°C} \cdot -4.1 - °C$$

$$Q_W = -1372.35 - J$$

Q= ?

m=80-g

C=4.184-J/g °C

ΔT= -4.1-°C

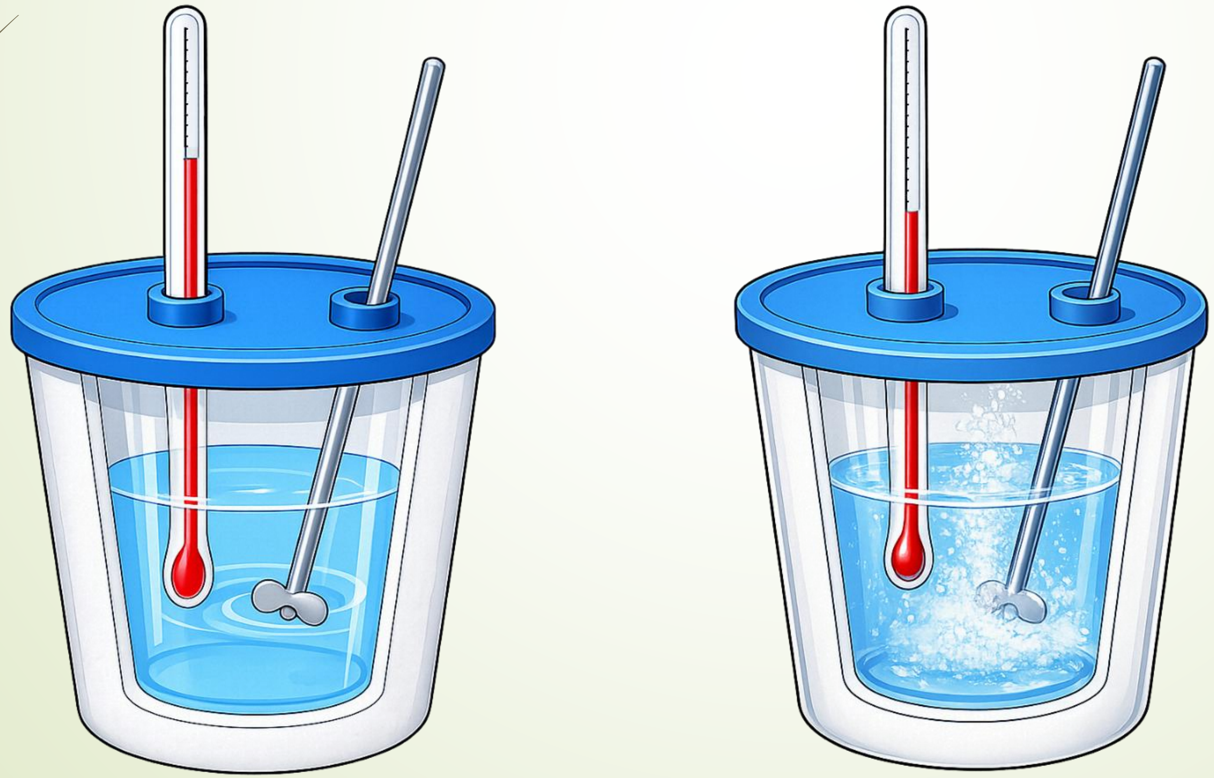
This is only the first part of the process.

Calorimetry From Partial to Full Solve

- To solve for the ΔH° of any salt:
  1. Use $Q = cm\Delta T$ to solve for the Q water (unit will be in J)
  2. The Q water will equal - the Q salt. ÷ 1000 to get kJ.
  3. Find out how many moles of salt you have.
  4. Set up a ratio and solve. (ΔH° unit is kJ/mol)

Calorimetry Example II

- 4.25-g of NH_4NO_3 dissolves in 20.0-g of water. The temperature drops from 20.3-°C to 8.0-°C. Calculate the ΔH°.



Note: The water temp went down so the reaction is Endothermic (ΔH is '+')

Calorimetry Example II

- 4.25-g of NH_4NO_3 dissolves in 20.0-g of water. The temperature drops from 20.3-°C to 8.0-°C. Calculate the ΔH°.

$$Q_{Water} = Cm\Delta T$$

$$Q_W = 20-g \cdot 4.184-\frac{J}{g°C} \cdot -12.3-°C$$

$$Q_W = -1029.264-J$$

Q_{Water} = ?

m= 20.0-g

C= 4.184-J/g·°C

ΔT= -12.3-°C

Calorimetry Example II

- If the water lost 1029.26-J it had to be from the NH_4NO_3 absorbing it.

2. Solve for the Q of the chemical (reaction).

Since enthalpy is given in units of kJ now would also be a good time to convert. (divide by 1000)

$$Q_{Salt} = -Q_{Water}$$

$$Q_{Salt} = -(-1029.264 - J)\ x\ \frac{k}{1000}$$

$$Q_{Salt} = 1.029264 - kJ$$

Calorimetry Example II

- Q_{Salt}=1.029-kJ is the heat absorbed due to 4.25-g of NH_4NO_3.
  3. Convert to moles.
  4. Set up a ratio.

$$4.25 - g\ NH_4NO_3 x\ \frac{1 - mol}{80.0 - g} = .0531 - mol$$

$$\Delta H° = \frac{Q_{Salt}}{n_{Salt}}$$

$$\Delta H° = \frac{1.029264 - kJ}{.0531 - mol} = 19.37$$

$\boxed{\Delta H° = 19.4\text{-kJ/mol}}$

Calorimetry Example III

- 56.0-g of H_3BO_3 dissolves in 50.0-g of water. The temperature rises from 15.0-°C to 33.4-°C. Calculate ΔH°.

$$Q_{Water} = Cm\Delta T$$

$$Q_W = 50 - g \cdot 4.184 - \frac{J}{g°C} \cdot 18.4 - °C$$

$$Q_W = 3849.28 - J$$

$$Q_{Salt} = -Q_{Water}$$

$$Q_{Salt} = -3.84928 - kJ$$

Q= ?

m=50.0-g

C=4.184-J/g·°C

ΔT= 18.4-°C

Calorimetry Example III

- H_3BO_3: Q_{salt}= -3.849-kJ m_{salt}= 56.0-g of.

$$n_{Salt} = 56-g\ H_3BO_3 x\ \frac{1-mol}{61.8-g} = .9061-mol$$

$$\Delta H° = \frac{Q_{Salt}}{n_{Salt}}$$

$$\Delta H° = \frac{-3.84928-kJ}{.9061-mol} = -4.24$$

$\boxed{\Delta H° = -4.2\text{-kJ/mol}}$