Unit 6

- 9 topics
- 7-9% of AP test
- Tentative test date: January 22

Today

- Simulation
- 6.1-6.3

6.1 Endothermic and Exothermic Processes

ENDURING UNDERSTANDING:

ENE-2	Changes in a substance's properties or change into a different substance requires an exchange of energy.

LEARNING OBJECTIVE:

ENE-2.A	Explain the relationship between experimental observations and energy changes associated with a chemical or physical transformation.

Enthalpy

All sorts of processes, both physical and chemical, have an **enthalpy** change associated with them.

While a general change in enthalpy is symbolised as ΔH, specific types of enthalpy changes may be symbolized by a subscripted explanation attached to this symbol.

By definition: $\Delta H = H_{products} - H_{reactants}$

$-\Delta H$ = exothermic (energy released to surroundings)

$+\Delta H$ = endothermic (energy absorbed from surroundings)

EXAMPLES:

$$KOH(s) \rightarrow K^+(aq) + OH^-(aq) \qquad \Delta H_{\text{solution}} = -57.8 \text{ kJ/mol}$$

$$C_3H_8(g) + 5\,O_2(g) \rightarrow 3\,CO_2(g) + 4\,H_2O(l) \qquad \Delta H_{\text{combustion}} = -2221 \text{ kJ/mol}$$

$$H_2O(s) \rightarrow H_2O(l) \qquad \Delta H_{\text{fusion}} = 6.0 \text{ kJ/mol}$$

$$Fe_2O_3(s) + 2\,Al(s) \rightarrow Al_2O_3(s) + 2\,Fe(s) \qquad \Delta H_{\text{reaction}} = -852 \text{ kJ/mol}$$

$$Ca(s) + O_2(g) + H_2(g) \rightarrow Ca(OH)_2(s) \qquad \Delta H_{\text{formation}} = -986 \text{ kJ/mol}$$

DEFINITIONS

System – the part of the world that you are studying. This might be a beaker, a salt, water and a thermometer or it could be just the salt that is dissolving.

Surroundings – everything else around the system. This might be the air, the tabletop, hands holding the beaker, or it might be the water and the beaker and the thermometer etc.

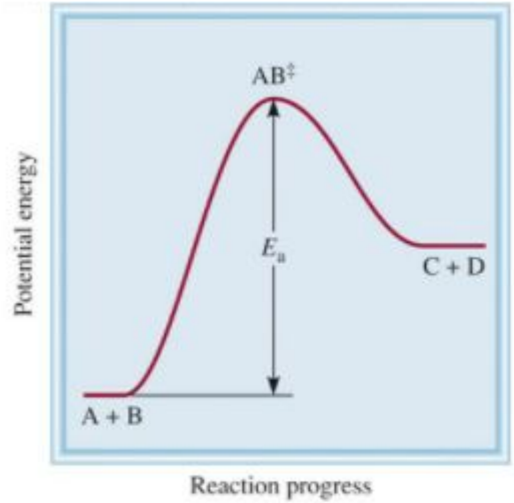
Universe - the system and the surroundings together.

Heat (q) - is the amount of thermal energy transferred

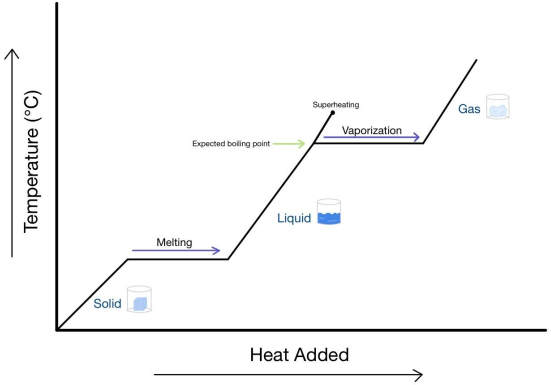
Work - the energy needed to move something against a force, in chemistry we generally think about expanding gases that push on a piston. Chemical work is often thought of as **work = $-P\Delta V$**

Endothermic

- Heat is absorbed by the system from the surroundings
- Enthalpy (ΔH) is positive
- The products are higher in energy than the reactants.
- Phase changes: melting, boiling, vaporization, sublimation
- Changes that involve breaking/separation.

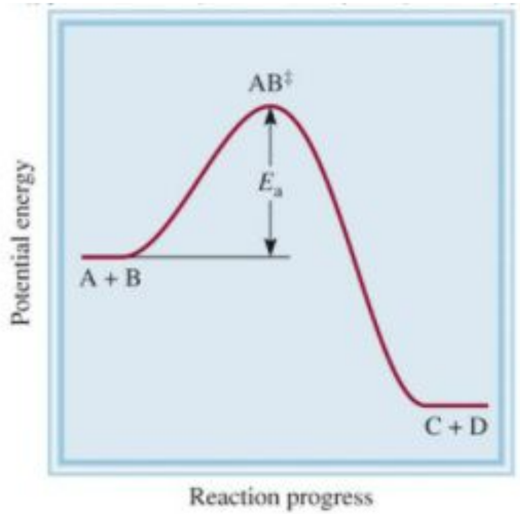


Phase Change Diagram

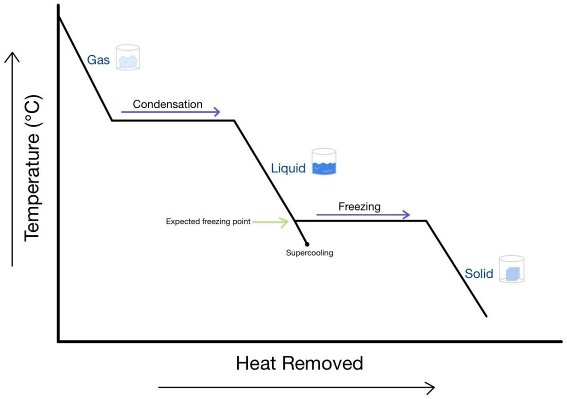


EXOTHERMIC

- Heat is released from the system to the surroundings
- Enthalpy (ΔH) is negative
- The products are lower in energy than the reactants.
- Phase changes: freezing, condensation,
- deposition
- Changes that involve combining materials.



Phase Change Diagram



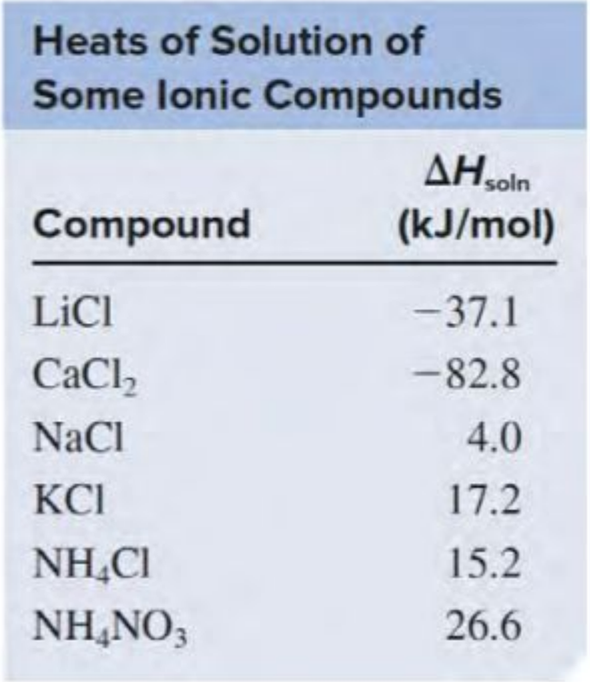
DISSOLVING CAN BE ENDOTHERMIC OR EXOTHERMIC

1) The solvent must expand by overcoming its intermolecular forces. This is an endothermic process.

2) The solute must expand by overcoming its intermolecular forces. This is an endothermic process.

3) The solute and solvent recombine, this process is exothermic.

*When we measure the temperature change of the solution, we are measuring the temperature change of the **surroundings** (not the system).



Heats of Solution of Some Ionic Compounds

Compound	ΔH_{soln} (kJ/mol)
LiCl	−37.1
$CaCl_2$	−82.8
NaCl	4.0
KCl	17.2
NH_4Cl	15.2
NH_4NO_3	26.6

Quick Lab

1. Record initial temperature of 50 mL water.

Is the water the system or the surroundings?

2. Add ~2 grams of Ammonium Chloride to the water.
3. Record the final temperature of the water.

Is the ΔH positive or negative?

Is the dissolution of NH_4Cl endothermic or exothermic?

4. Repeat steps with Copper Chloride.

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\begin{aligned} \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ \\ &= -RT \ln K \\ &= -nFE^\circ \end{aligned}$$

$$I = \frac{q}{t}$$

$$E_{cell} = E_{cell}^\circ - \frac{RT}{nF} \ln Q$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)
Q = reaction quotient

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

Together

Instant coffee is made through a process called freeze drying. First the coffee is made in the traditional way, and then the coffee is cooled until it becomes a solid, finally the pressure is decreased and the ice changes to a gas, leaving behind the freeze-dried coffee granules.

Identify the two phase changes that take place and classify them as endothermic or exothermic.

YOU DO, WE REVIEW

A solution of ammonium nitrate was created by dissolving 5.02 grams of ammonium nitrate in 100.0 mL of water at 22.3° C. After forming the solution the temperature was 17.3°C.

Did heat enter the system or leave the system?

What is the sign for q?

Was the dissolution process endothermic or exothermic?

6.2 Energy Diagrams & 6.3 Heat Transfer and Thermal Equilibrium

ENDURING UNDERSTANDING:

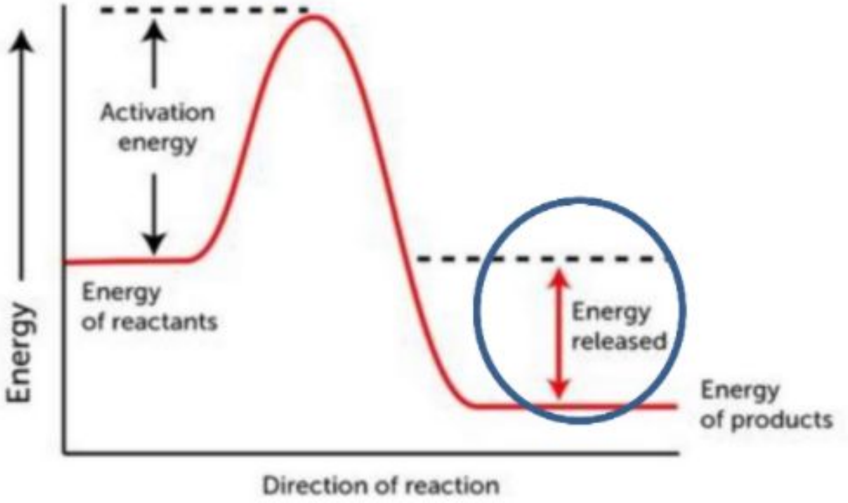
ENE-2	Changes in a substance’s properties or change into a different substance requires an exchange of energy.

LEARNING OBJECTIVES:

ENE-2.B	Represent a chemical or physical transformation with an energy diagram.
ENE-2.C	Explain the relationship between the transfer of thermal energy and molecular collisions.

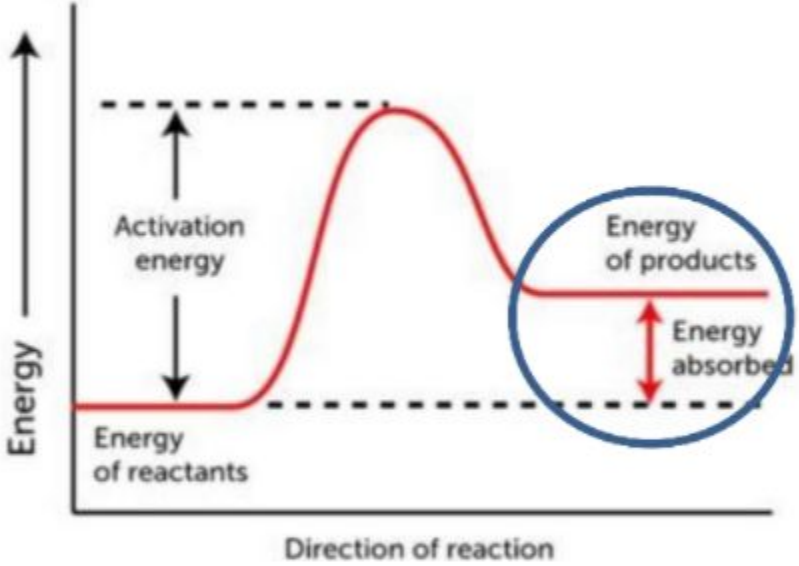
EXOTHERMIC REACTIONS

- Energy is released from the system into the surroundings, Heat EXITS (leaves the system)
- The products are lower in energy than the reactants (because energy has been released)



ENDOTHERMIC REACTIONS

- Energy is absorbed by the system from the surroundings, Heat ENTERS (goes into the system)
- The products are higher in energy than the reactants (because energy has been absorbed)



ENTHALPY OF REACTION

The difference in the energy of the products and energy of the reactants is called the **enthalpy of reaction**, or **heat of reaction**, and is represented as $\Delta \mathbf{H}_{\mathbf{rxn}}$ in $\text{kJ/mol}_{\text{rxn}}$.

$$\Delta H^\circ_{\text{reaction}} = \Sigma \Delta Hf\,^\circ_{\text{products}} - \Sigma \Delta Hf\,^\circ_{\text{reactants}}$$

The sign for ΔH is:

- Negative for exothermic reactions
- Positive for endothermic reactions
- ΔH is labeled as the difference between the energies of the products and reactants.

THERMODYNAMICS/ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\begin{aligned}\Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ \\ &= -RT \ln K \\ &= -nFE^\circ\end{aligned}$$

$$I = \frac{q}{t}$$

$$E_{cell} = E_{cell}^\circ - \frac{RT}{nF} \ln Q$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)
Q = reaction quotient
Faraday's constant, F = 96,485 coulombs per mole of electrons

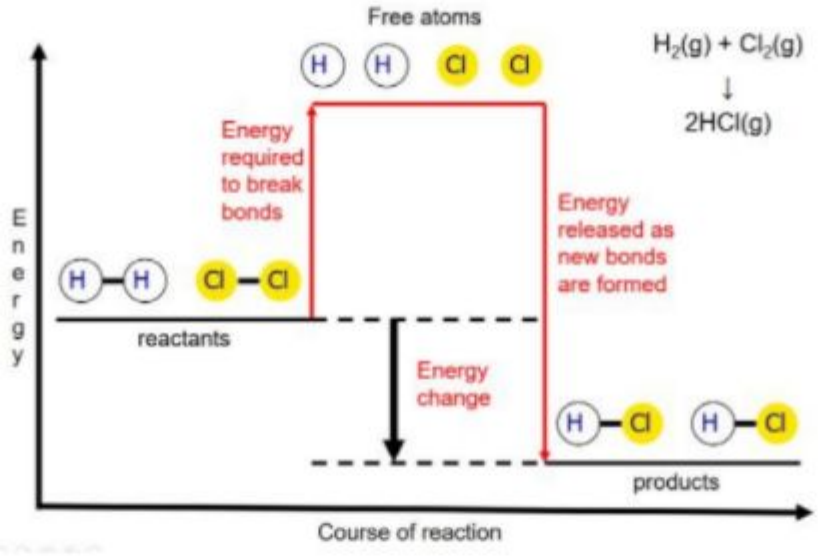
$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

ACTIVATION ENERGY

The activation energy is used to break bonds and/or intermolecular forces

Energy released from the reaction comes from forming new bonds/attractions.

If more energy is released then the reaction is exothermic overall, and if more energy is required then the reaction is endothermic overall.



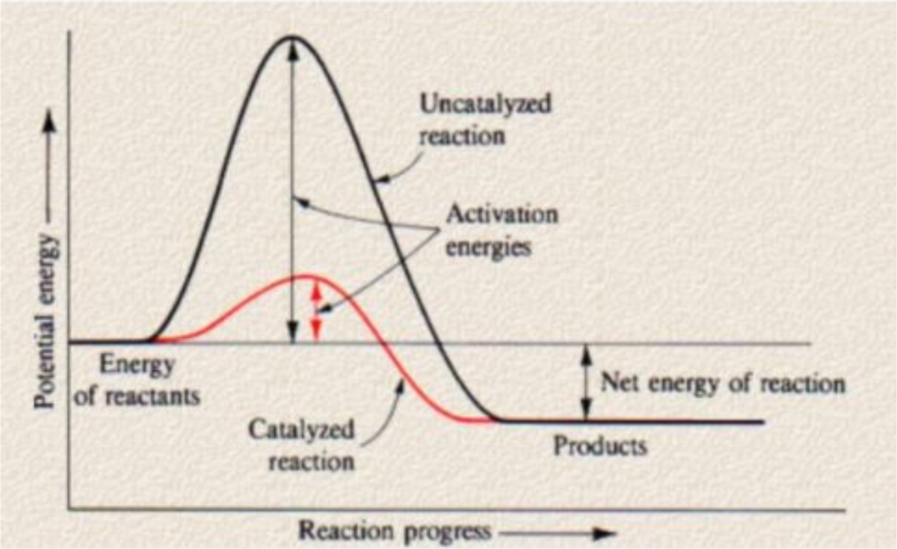
CATALYST

Catalysts are added to reactions to increase the reaction rate.

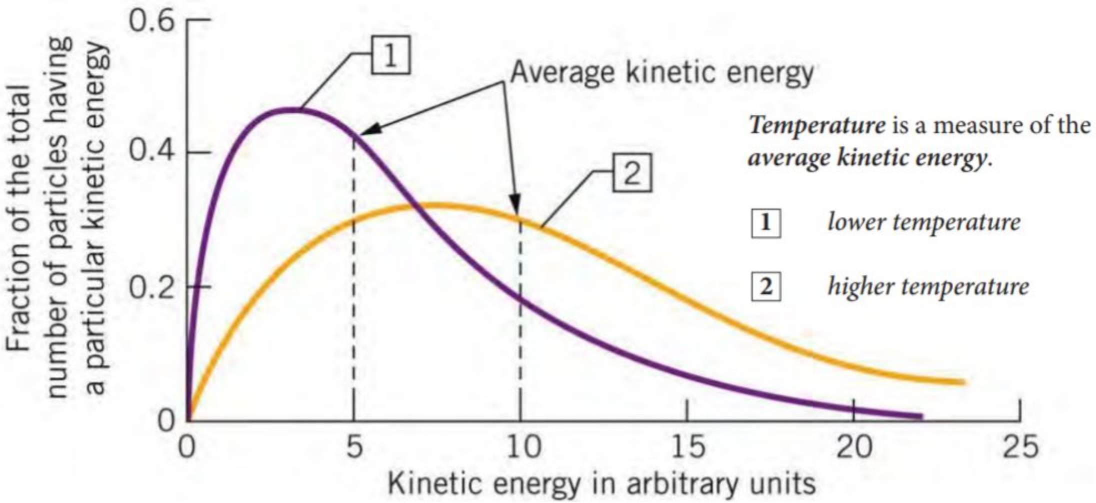
They are not consumed in the reaction.

Some catalysts provide an alternate pathway with lower activation energies.

ΔH is NOT changed by a catalyst!!



Maxwell-Boltzmann Distribution Curves



In Unit 3 we saw the Maxwell-Boltzmann distribution curves for substances at different temperatures.

The particles with a higher temperature have greater kinetic energy than those that are at a lower temperature.

Temperature is a measure of the average kinetic energy of the particles of the sample.

KE = ½ mv2. (m = mass, v = velocity)

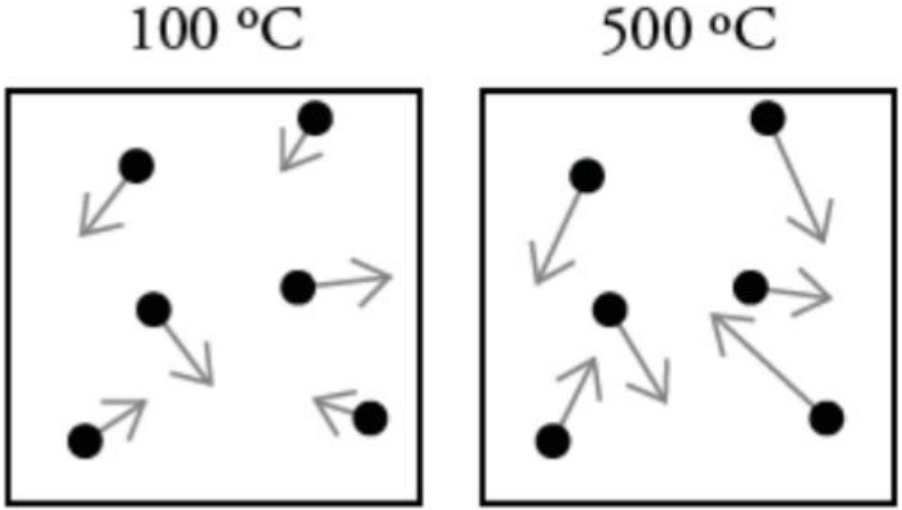
PARTICLE DIAGRAMS

At any given temperature, particles will have a range of kinetic energies and a range of speeds.

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

Some particles will be low speed, most will be close to average speed and a few will have high speed.

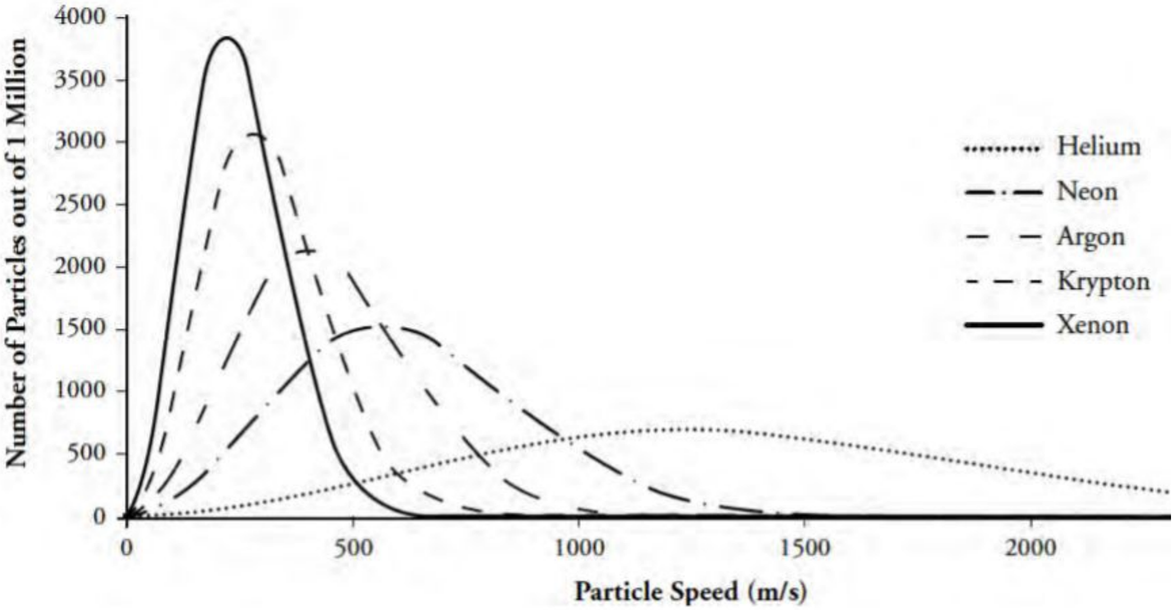
They all share the same average speed.



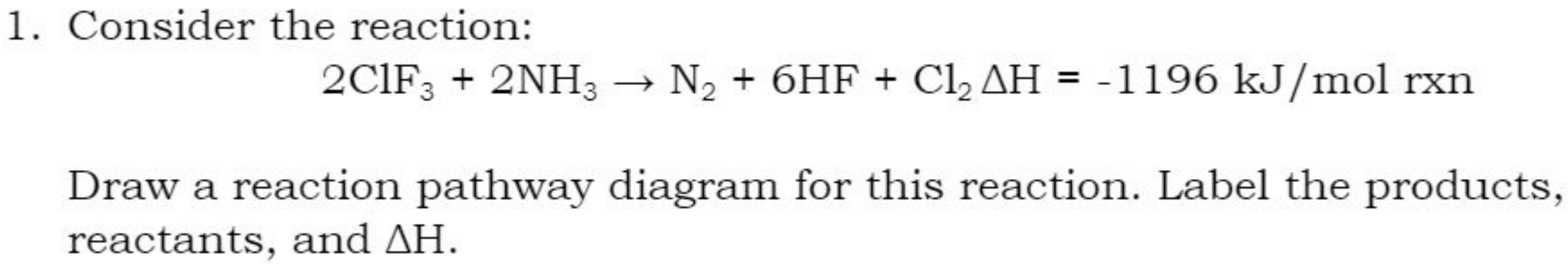
Don't forget about Mass!

Different substances at the same temperature have the same average Kinetic Energy.

Differences between the substances - in particular particle mass, means that their average speeds will be different.



TOGETHER 6.2



1. Consider the reaction:

$$2ClF_3 + 2NH_3 \rightarrow N_2 + 6HF + Cl_2 \quad \Delta H = -1196\ \text{kJ/mol rxn}$$

Draw a reaction pathway diagram for this reaction. Label the products, reactants, and ΔH.

HEAT TRANSFER

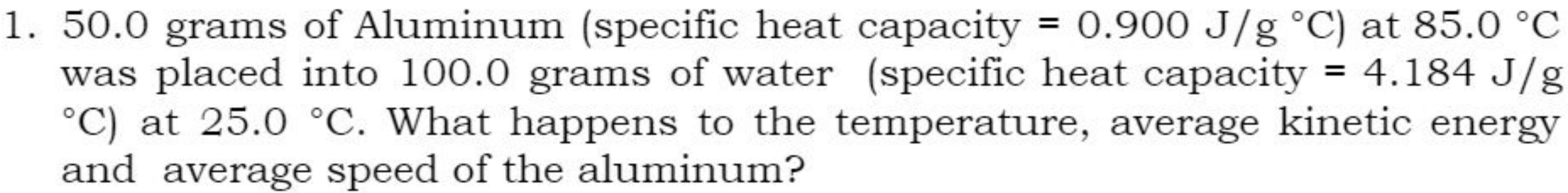
When hotter particles and colder particles are in contact with one another and they collide, energy is transferred between them. Names for this are: **heat transfer** or **heat exchange** or **transfer of energy as heat**.

This results in the hotter particles losing some energy and the colder substance increases in temperature. Heat is always transferred from the hotter substance to the colder one.

If the two substances result in the same temperature, we call that **thermal equilibrium**.

When two substances reach the same temperature, they have the same average kinetic energy.

TOGETHER 6.3



1. 50.0 grams of Aluminum (specific heat capacity = 0.900 J/g °C) at 85.0 °C was placed into 100.0 grams of water (specific heat capacity = 4.184 J/g °C) at 25.0 °C. What happens to the temperature, average kinetic energy and average speed of the aluminum?

6.4 Heat Capacity and Calorimetry

ENDURING UNDERSTANDING:

ENE-2	Changes in a substance’s properties or change into a different substance requires an exchange of energy.

LEARNING OBJECTIVE:

ENE-2.D	Calculate the heat q absorbed or released by a system undergoing heating/ cooling based on the amount of the substance, the heat capacity, and the change in temperature.

THERMODYNAMICS/ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

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

$$= -RT \ln K$$

$$= -nFE^\circ$$

$$I = \frac{q}{t}$$

$$E_{cell} = E^\circ_{cell} - \frac{RT}{nF}\ln Q$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)
Q = reaction quotient
Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

HEAT (Q)

Definition: A form of energy that flows between two samples of matter because of their differences in temperature. Label: q

Units: Joules or calories

Quantity of heat (q) is the transfer of thermal energy between two bodies that are at different temperatures, from an object at a higher temperature to one that is colder.

There are three factors that contribute to the amount of heat (q) transferred:

- mass of the object (m)
- specific heat capacity (c)
- change in temperature ($\Delta T = T_{final} - T_{initial}$).

SPECIFIC HEAT CAPACITY

Definition: Amount of energy needed to raise the temperature of one gram of a substance by one ° Celsius (or 1 Kelvin).

Label: c

Units: joules/gram °Celsius (or Kelvin)

Each substance has its own specific heat capacity.

Examples:

Water: 4.184 J/g °C

Air: 1.005 J/g °C

Lead: 0.129 J/g °C (*Metals tend to have low specific heats)

QUICK CHECK

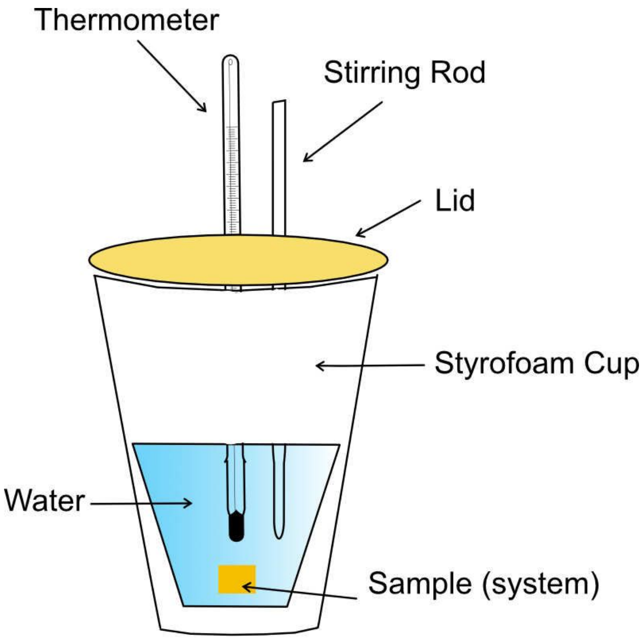
1. How do the specific heat capacities of metals compare to that of liquid water?
2. If equal masses of water and iron were exposed to the same heat source, which one would reach the temperature of 100°C first?
3. Consider equal masses of iron and water at 100°C. Which would cool to room temperature more quickly?

CALORIMETRY

A **calorimeter** measures the amount of heat transferred during a reaction. A calorimeter is essentially a device that stops the transfer of heat between the inside of the calorimeter and the surroundings.

Calorimetry is the measurement of the quantity of heat exchanged.

For example, the energy from an exothermic chemical reaction is absorbed in a container of water, the change in temperature of the water provides a measure of the amount of heat added.



HEAT CALCULATIONS

Heat calculations involving physical processes with only the system and the surroundings require a generalized version of the **first law of thermodynamics** - the idea that energy cannot be created or destroyed; energy must be conserved.

Heat Lost (by hotter object) = Heat Gained (by colder object)

The heat is transferred from the hot metal to the colder water until they reach thermal equilibrium. We can measure the heat gained by the water to find the heat lost by the metal.

$$q_{(metal)} = -q_{(water)}$$

We can Solve for unknowns!

$$q_{(metal)} = -q_{(water)}$$

Since we know that the formula to calculate heat when there is a temperature change is: $q = m\ C\ \Delta T$

We can substitute to form:

$$m_{(metal)}\ C_{(metal)}\ \Delta T_{(metal)} = -\ m_{(water)}\ C_{(water)}\ \Delta T_{(water)}$$

Different Specific Heat Capacity means different temp changes

Each substance has its own specific heat capacity.

Compare aluminum and lead, the specific heat of aluminum is 0.900 J/g °C while the specific heat of lead is 0.160 J/g °C. Adding 1000 J of energy to 100 grams of each substance results in different changes in temperature as shown below:

Aluminum	Lead
$q=mC\Delta T$ $\Delta T = q / mC$ $\Delta T = 1000\ \text{J}/((100\ \text{g})(0.900\ \text{J/g°C}))$ $\Delta T = 11.1\text{°C}$	$q=mC\Delta T$ $\Delta T = q / mC$ $\Delta T = 1000\ \text{J}/((100\ \text{g})(0.160\ \text{J/g°C}))$ $\Delta T = 62.5\text{°C}$

You can see that when the specific heat is higher, the change in temperature is lower and vice versa.

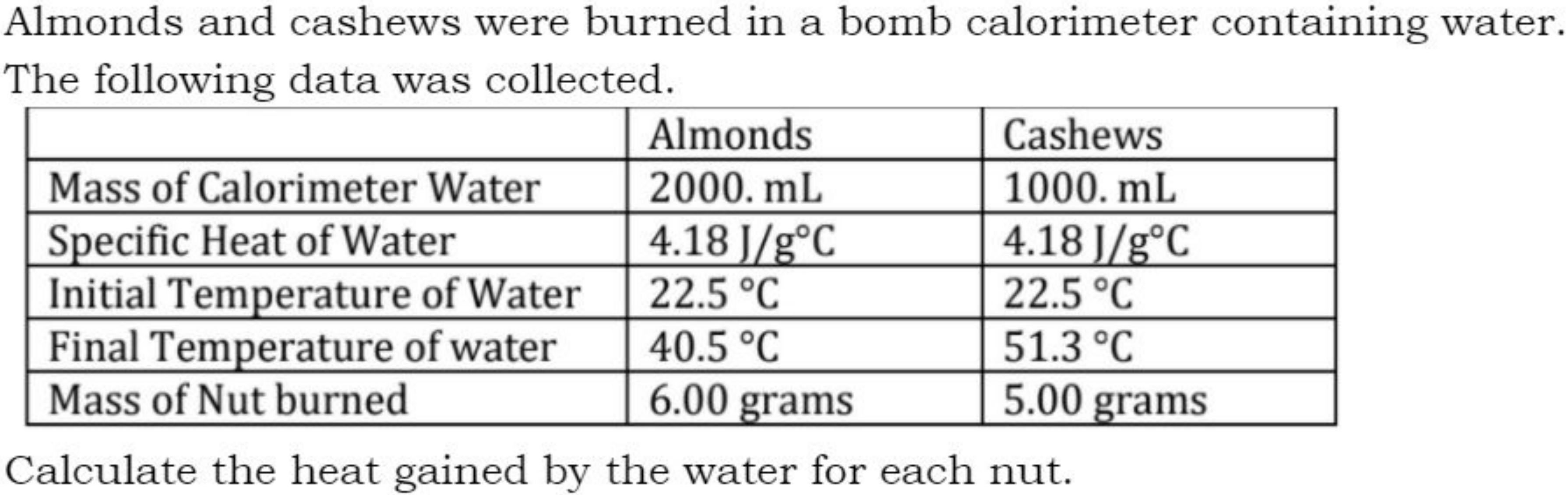
ENERGY CHANGES

Chemical systems change their energy through three main processes:

heating/cooling, **phase transitions**, and **chemical reactions**. We have seen that we can use $q=mC\Delta T$ to calculate the heat when a temperature change occurs, in other words when there is a change in the kinetic energy of the particles. This is shown as the areas with a positive slope on the heating of water curve above.

In the next lesson we will learn how to calculate the heat when phase transition takes place. During phase changes, there is no change in temperature; the energy is used to change the position of the particles relative to one another, in other words the potential energy is changing. When a phase change occurs we need to use a different formula to calculate the heat.

TOGETHER



Almonds and cashews were burned in a bomb calorimeter containing water. The following data was collected.

	Almonds	Cashews
Mass of Calorimeter Water	2000. mL	1000. mL
Specific Heat of Water	4.18 J/g°C	4.18 J/g°C
Initial Temperature of Water	22.5 °C	22.5 °C
Final Temperature of water	40.5 °C	51.3 °C
Mass of Nut burned	6.00 grams	5.00 grams

Calculate the heat gained by the water for each nut.

YOU DO, WE REVIEW

50.0 mL of 0.500 M HCl was added to 50.0 mL of 0.500 M NaOH, the initial temperature of the solutions was 19.8 °C. The reaction below occurred:

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

The final temperature of the mixture was 26.3°C.
Calculate the heat gained by the mixture.

What is the heat of reaction per mole of NaOH?

AGENDA

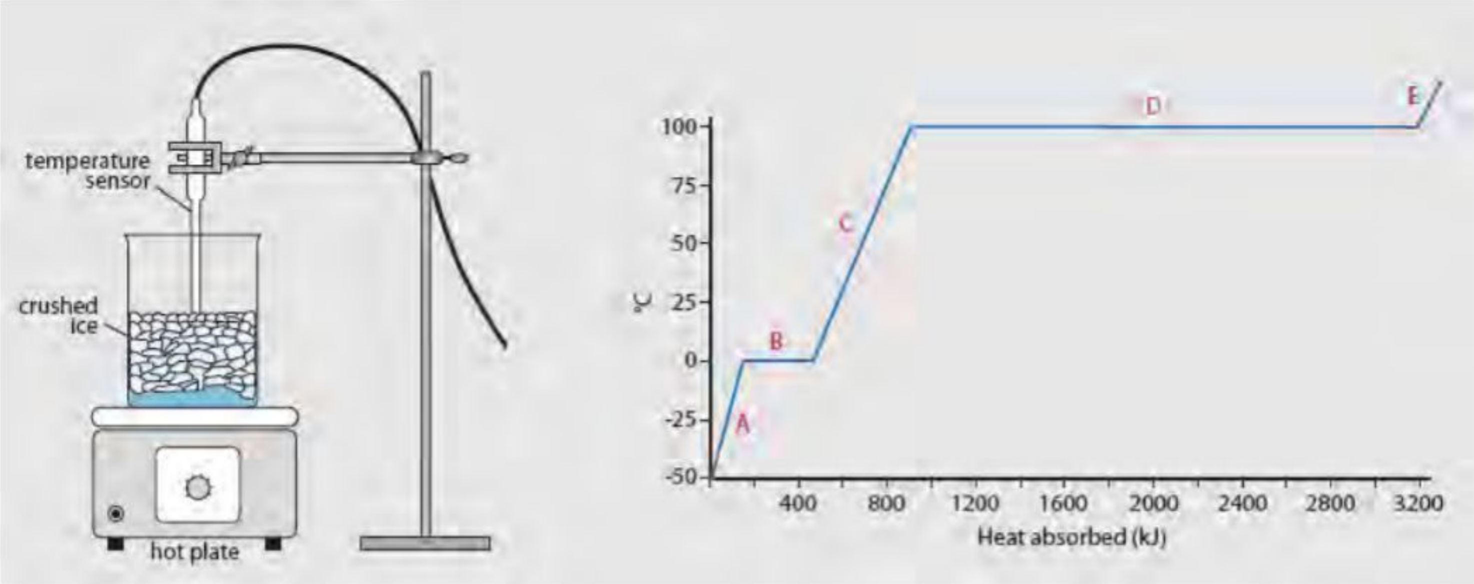
- How does a rice cooker work
- 6.5 Notes, Worksheet, Packet problems
- 6.6 Notes, Packet problems

- Homework: Finish 6.5 and 6.6 packet problems

6.5 Energy of Phase Changes

ENDURING UNDERSTANDING:	
ENE-2	Changes in a substance's properties or change into a different substance requires an exchange of energy.
LEARNING OBJECTIVE:	
ENE-2.E	Explain changes in the heat q absorbed or released by a system undergoing a phase transition based on the amount of the substance in moles and the molar enthalpy of the phase transition.

Fill in what is happening in each region and approx. how much energy is required.



Region	Energy is used to...	Energy required
A	Heat solid water from −50 to 0°C	Approximately 180 kJ of energy
B		
C		
D		
E		

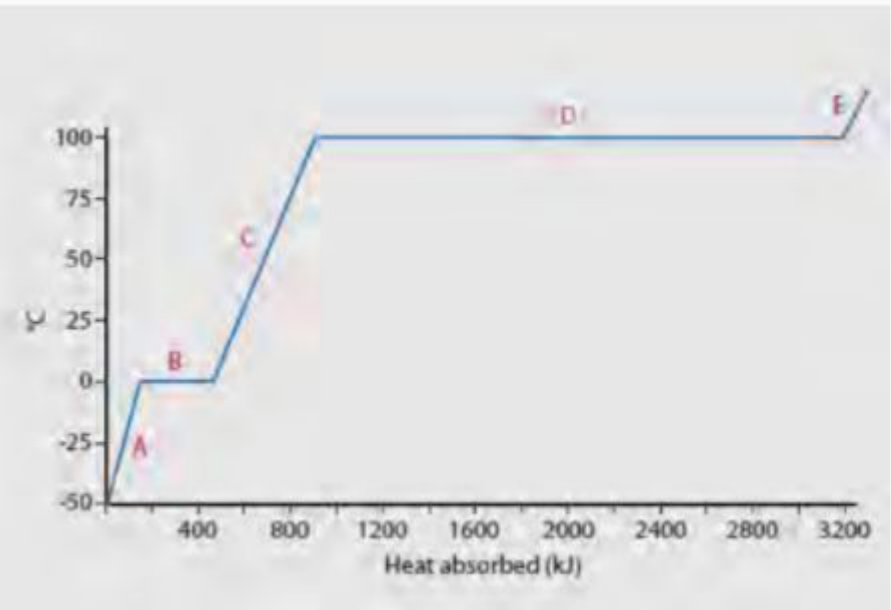
THINK ABOUT IT:

Regions A, C, and E:

1. What is happening to the temperature?
2. What is happening to the average speed of the H_2O molecules?
3. What is the heat energy supplied doing?

Regions B and D:

1. What two things might cause the energy required to differ?
2. What is happening to the temperature?
3. What is happening to the average speed of the H_2O molecules? A
4. What is the heat energy supplied doing?



Two Main Effects of Heat Energy Supplied

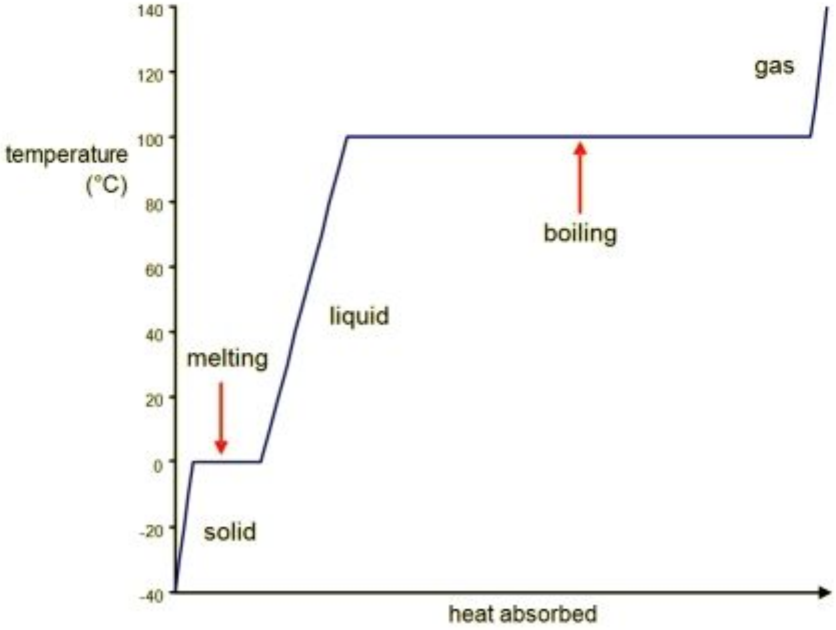
1. to **increase the Kinetic Energy** (translation, vibration and/or rotation) which can increase average speed and **increases temperature**. This is governed by the specific heat capacity of the material and effects can be calculated using: $q = c\ m\ \Delta T$ or $q = c\ n\ \Delta T$
2. **- to Overcome IMFs** (move particles further apart) which has **no effect on average speed and no effect on temperature**. This is often referred to as latent (hidden) heat due to the fact that there is no obvious effect (no temperature rise) associated with this heat.

PHASE CHANGES

When heat is released or absorbed during phase transitions there is no change in temperature; the energy is used to change the position of the particles relative to one another.

This is shown as the plateaus on the heating or cooling curve. Notice that the segment for the boiling is longer than the segment for melting, this is because separating liquid particles to form a gas requires more energy than separating molecules from their solid state to a liquid state.

Potential energy is changing.



CALCULATING Q

To calculate the heat needed to perform a phase change occurs we use the enthalpy for that phase change and multiply by the amount of material.

$q = m$ (or n) ΔH

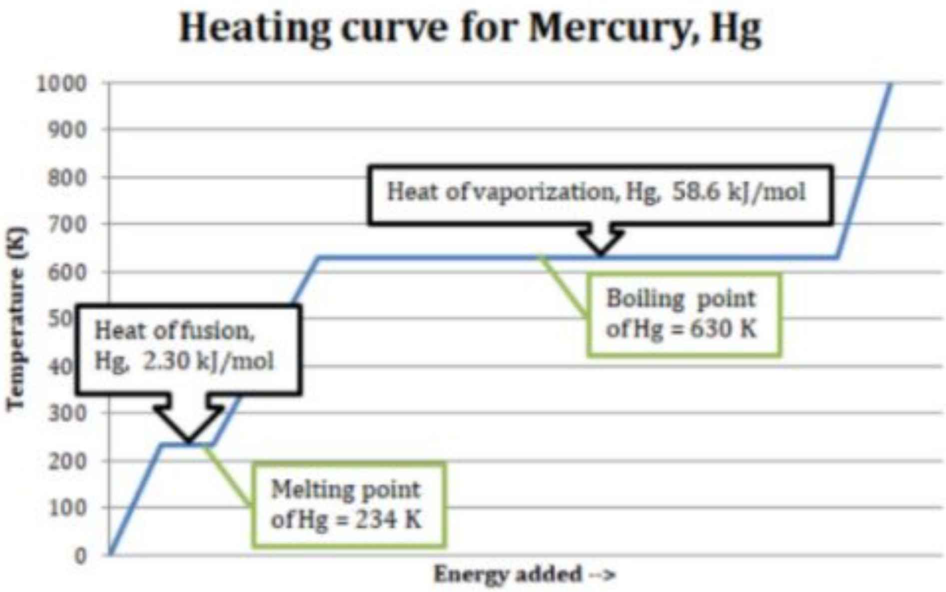
q = Heat

m = mass (or n = moles)

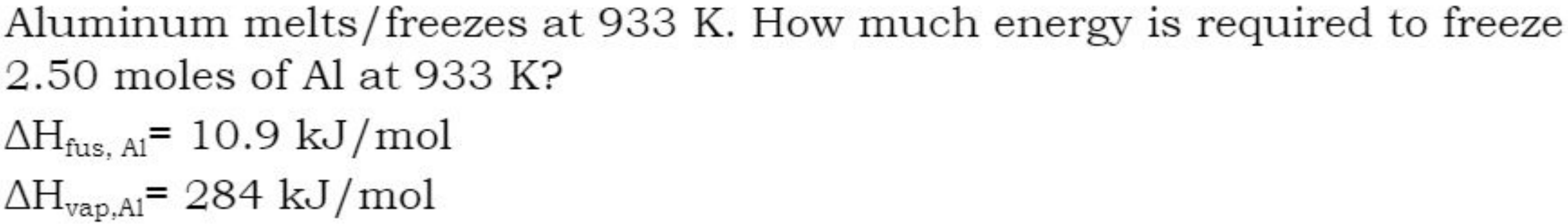
ΔH = enthalpy for the phase change “Heat of ___” or “Molar heat of ___”

Together

Given the heating curve for mercury, calculate the heat to freeze 100.0 grams of mercury, Hg, at its melting point.



YOU DO, WE REVIEW



Aluminum melts/freezes at 933 K. How much energy is required to freeze 2.50 moles of Al at 933 K?

$\Delta H_{fus,\ Al}$= 10.9 kJ/mol

$\Delta H_{vap,Al}$= 284 kJ/mol

6.6 Introduction to Enthalpy of Reaction

ENDURING UNDERSTANDING:

ENE-2	Changes in a substance’s properties or change into a different substance requires an exchange of energy.

LEARNING OBJECTIVE:

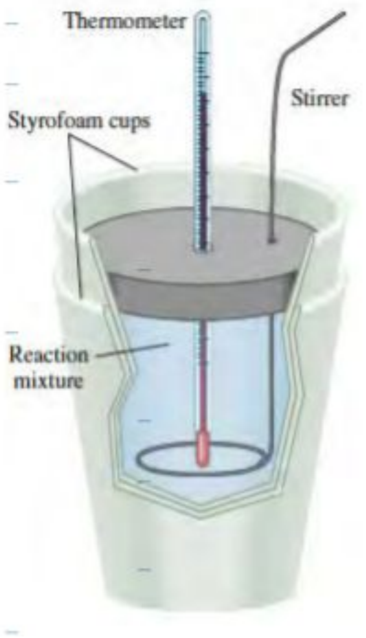
ENE-2.F	Calculate the heat q absorbed or released by a system undergoing a chemical reaction in relationship to the amount of the reacting substance in moles and the molar enthalpy of reaction.

Two Types of Calorimetry

In the laboratory, heat changes in physical and chemical processes are measured with a calorimeter, a closed container designed specifically for this purpose.

A **constant-pressure calorimeter** made of two Styrofoam coffee cups.

- The outer cup helps to insulate the reacting mixture from the surroundings.
- Two solutions of known volume containing the reactants at the same temperature are carefully mixed in the calorimeter.
- The heat produced or absorbed by the reaction can be determined by measuring the temperature change.



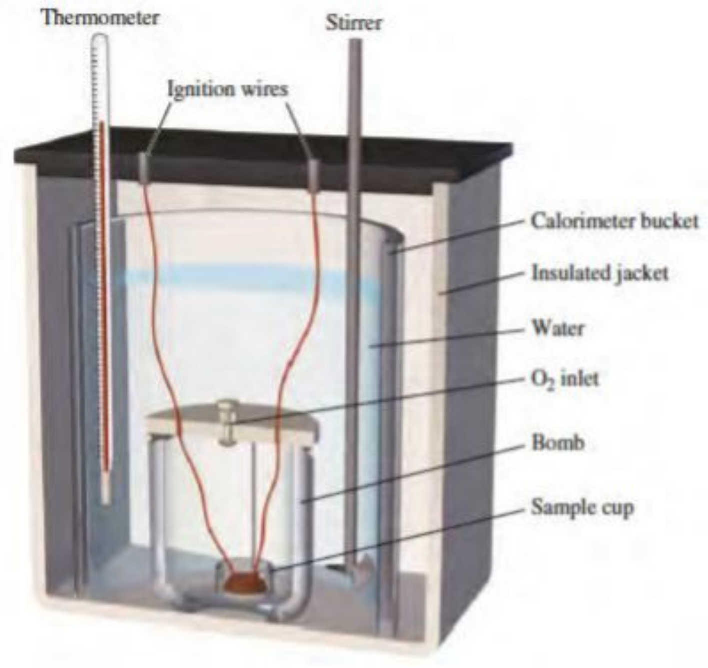
Two Types of Calorimetry

A **constant-volume bomb calorimeter:**

The calorimeter is filled with oxygen gas before it is placed in the bucket.

The sample is ignited electrically, and the heat produced by the reaction can be accurately determined by measuring the temperature increase in the known amount of surrounding water.

This is the type of apparatus routinely used to determine the calorific value of many foods. 1 Cal = 1 kcal = 4.184 kJ.



HEAT GAIN OR LOSS REPRESENTATION

$$N_{2\,(g)} + O_{2\,(g)} \rightarrow 2\ NO_{(g)} \quad \Delta H = +\ 180\ \text{kJ/mol}\ _{\text{rxn}}$$

$$N_{2\,(g)} + O_{2\,(g)} + 180\ \text{kJ} \rightarrow 2\ NO_{(g)}$$

Both reactions indicate that the reaction is endothermic; heat is absorbed in the course of the reaction.

The reaction shows the amount of heat energy required for the reaction as it is written; one mole of each reactant to form two moles of the product.

EXAMPLE 1

$$N_{2\,(g)} + O_{2\,(g)} \rightarrow 2\,NO_{\,(g)} \quad \Delta H = +\,180\ \mathrm{kJ/mol\ rxn}$$

$$N_{2\,(g)} + O_{2\,(g)} + 180\ \mathrm{kJ} \rightarrow 2\,NO_{\,(g)}$$

How much heat (q) is required to react 2.4 moles of N_2 (g) with excess O_2 (g) to form NO?

Example 2

$$N_{2\,(g)} + O_{2\,(g)} \rightarrow 2\,NO_{\,(g)} \quad \Delta H = +\,180\ \text{kJ/mol}\ \text{rxn}$$

$$N_{2\,(g)} + O_{2\,(g)} + 180\ \text{kJ} \rightarrow 2\,NO_{\,(g)}$$

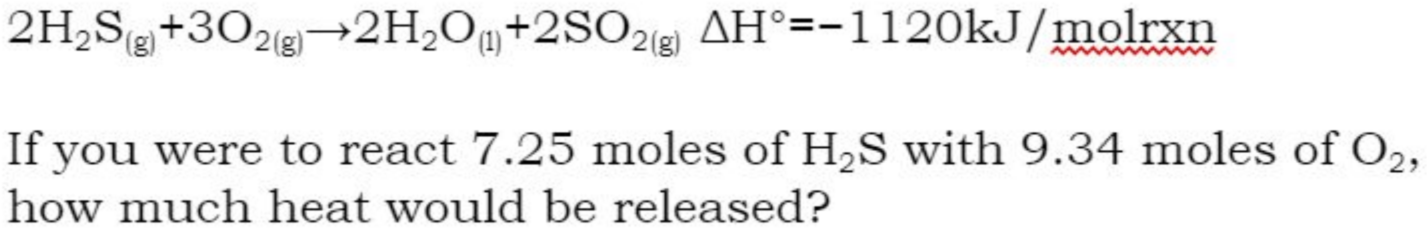
If 25 kJ of heat was absorbed in the reaction, how many moles of NO were produced?

AT CONSTANT PRESSURE, $\Delta H = Q$

*common for most AP reactions

- $\Delta E_{sys} = q + w$ (First law of thermodynamics)
- $\Delta H_{sys} = \Delta E_{sys} + \Delta(PV)$
- $w = -P\Delta V$
- $\Delta H_{sys} = q_p + w + P\Delta V$ (at constant pressure)
- $\Delta H_{sys} = q_p + -P\Delta V + P\Delta V$
- $\boldsymbol{\Delta H_{sys} = q_p}$

Together



$2H_2S_{(g)}+3O_{2(g)}\rightarrow 2H_2O_{(l)}+2SO_{2(g)}$ $\Delta H^\circ=-1120\text{kJ/molrxn}$

If you were to react 7.25 moles of H_2S with 9.34 moles of O_2, how much heat would be released?

YOU DO, WE REVIEW

$2\ H_{2(g)} + O_{2\ (g)} \rightarrow 2\ H_2O\ _{(g)}$ $\Delta H° = -481.6$ kJ/molrxn

If you performed this reaction and produced 92.1 kJ of energy, how many moles of water were also produced?

Agenda

- Bellwork
- Homework check
- 6.7
  - Notes
  - Videos
  - Practice Problems
- 6.8
  - Notes
  - Videos
  - Practice Problems

6.7 Bond Enthalpies

ENDURING UNDERSTANDING:

ENE-3	The energy exchanged in a chemical transformation is required to break and form bonds.

LEARNING OBJECTIVE:

ENE-3.A	Calculate the enthalpy change of a reaction based on the average bond energies of bonds broken and formed in the reaction.

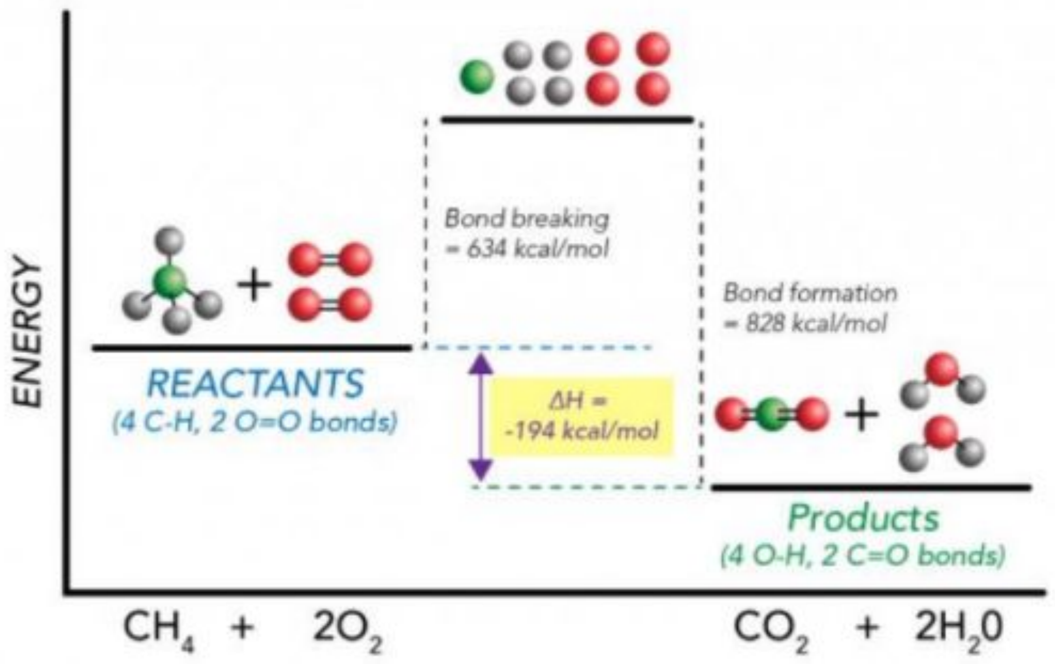
I can

- Calculate enthalpy of a reaction using the energy needed to break bonds and the energy released from bonds

BOND ENTHALPY

In a chemical reaction, the bonds in the reactants must be broken and remade into the products.

- Breaking bonds always requires energy
- Forming bonds always releases energy.

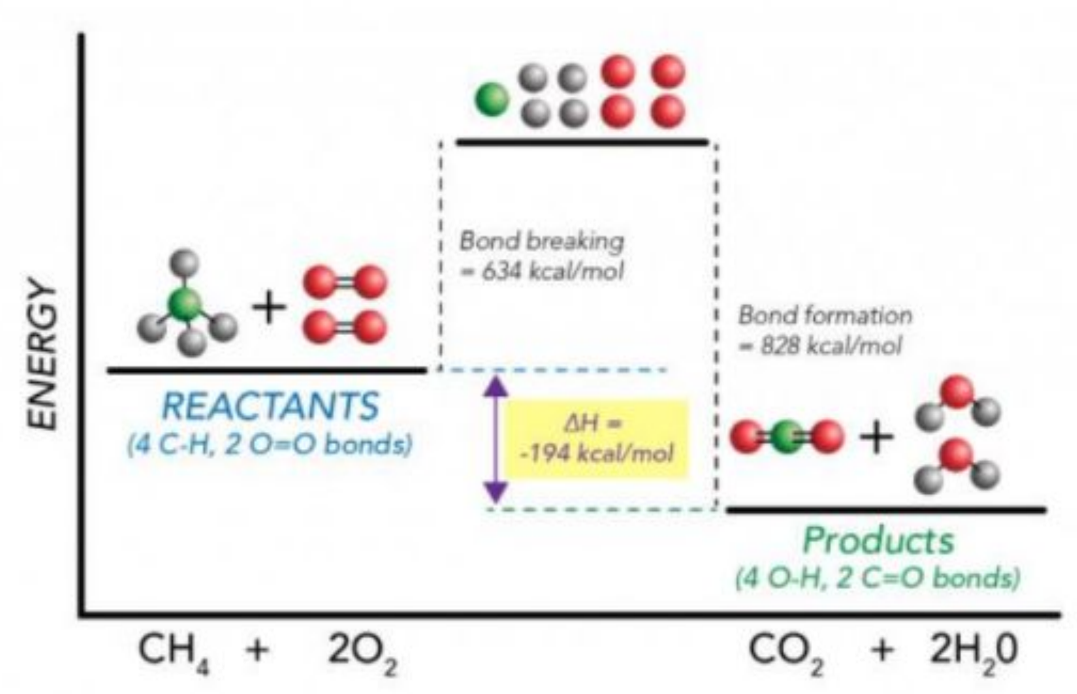


BOND ENTHALPY

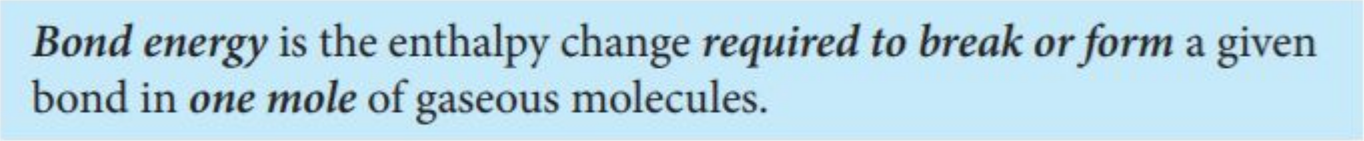
You can estimate the potential energy change for a reaction by adding up the bond energies for breaking and then adding up the bond energies for making

If the energy breaking requires more energy than making _____

If making releases more energy than breaking ______

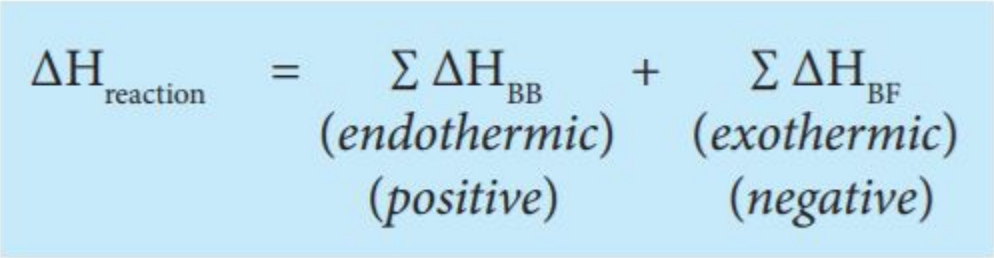


Bond Enthalpy



Bond energy is the enthalpy change ***required to break or form*** a given bond in ***one mole*** of gaseous molecules.

Bond enthalpies are often listed as POSITIVE values (bond breaking). When a bond is formed, make sure to put a NEGATIVE sign in front of that value.



$$\Delta H_{\text{reaction}} = \underset{\substack{(\textit{endothermic})\\(\textit{positive})}}{\Sigma\, \Delta H_{\text{BB}}} + \underset{\substack{(\textit{exothermic})\\(\textit{negative})}}{\Sigma\, \Delta H_{\text{BF}}}$$

Average Bond Energy (kJ mol^{-1})

$$CH_{4\,(g)} + 2O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2H_2O_{(g)}$$

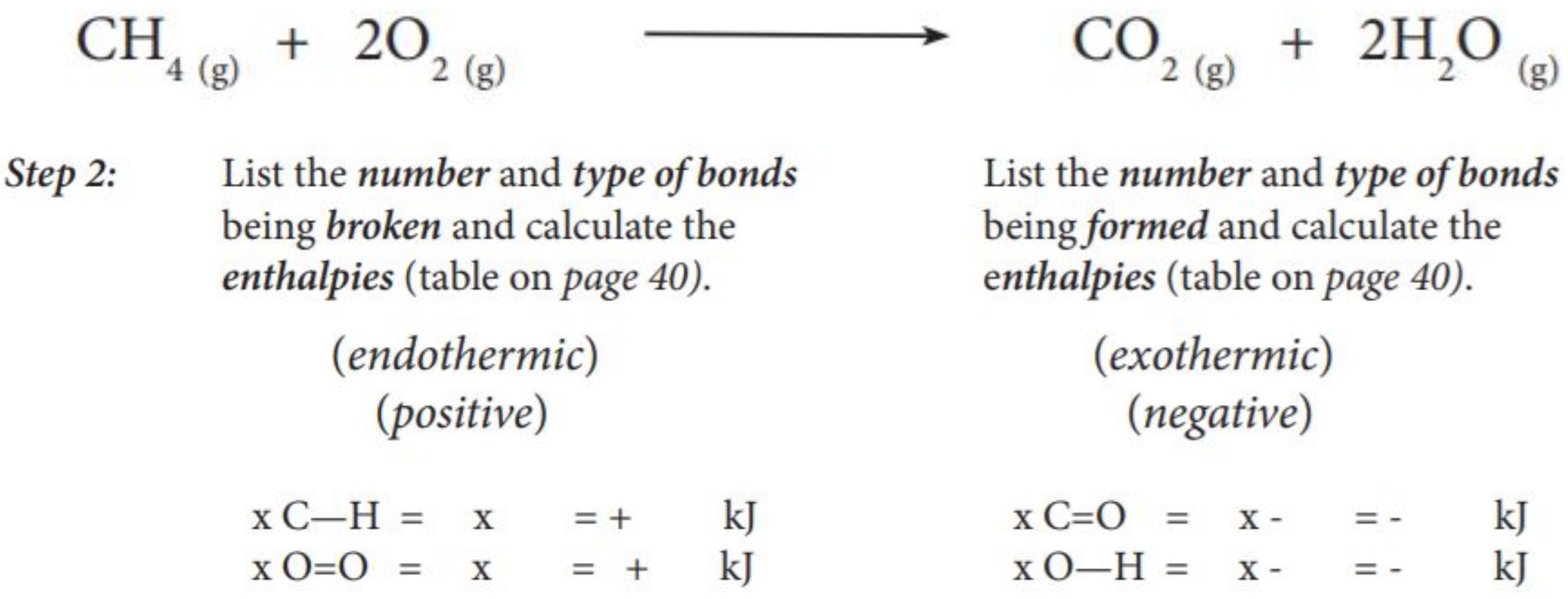
Bond	Energy	Bond	Energy	Bond	Energy	Bond	Energy
Single Bonds							
H–H	432	N–H	391	Si–H	323	S–H	347
		N–Cl	200	Si–Cl	381	S–I	~170
C–H	413	N–Br	243	Si–Br	310		
C–C	347	N–I	159	Si–I	234	F–F	159
C–Si	301					F–Cl	193
C–N	305	O–H	467	P–H	320	F–Br	212
C–O	358	O–P	351	P–Si	213	F–I	263
C–P	264	O–O	204	P–P	200	Cl-Cl	243
C–S	259	O–S	265	P–F	490	Cl-Br	215
C–F	453	O–F	190	P–Cl	331	Cl-I	208
C–Cl	339	O–Cl	203	P–Br	272	Br–Br	193
C–Br	276	O–Br	234	P–I	184	Br–I	175
C–I	216	O–I	234			I–I	151
Multiple Bonds							
C=C	614	N=N	418	C≡C	839	N≡N	945
C=N	615	N=O	607	C≡N	891		
C=O	745 (799 in CO_2)	O_2	498	C≡O	1070		

Walk Through example

$$CH_{4\,(g)} + 2O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2H_2O_{\,(g)}$$

Step 1: If necessary, quickly sketch structural formulae to insure you ***count all the bonds*** correctly and ***identify each type of bond*** correctly.

WALK THROUGH EXAMPLE



$$CH_{4\,(g)} + 2O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2H_2O_{(g)}$$

Step 2: List the ***number*** and ***type of bonds*** being ***broken*** and calculate the ***enthalpies*** (table on *page 40*).

(endothermic)
(positive)

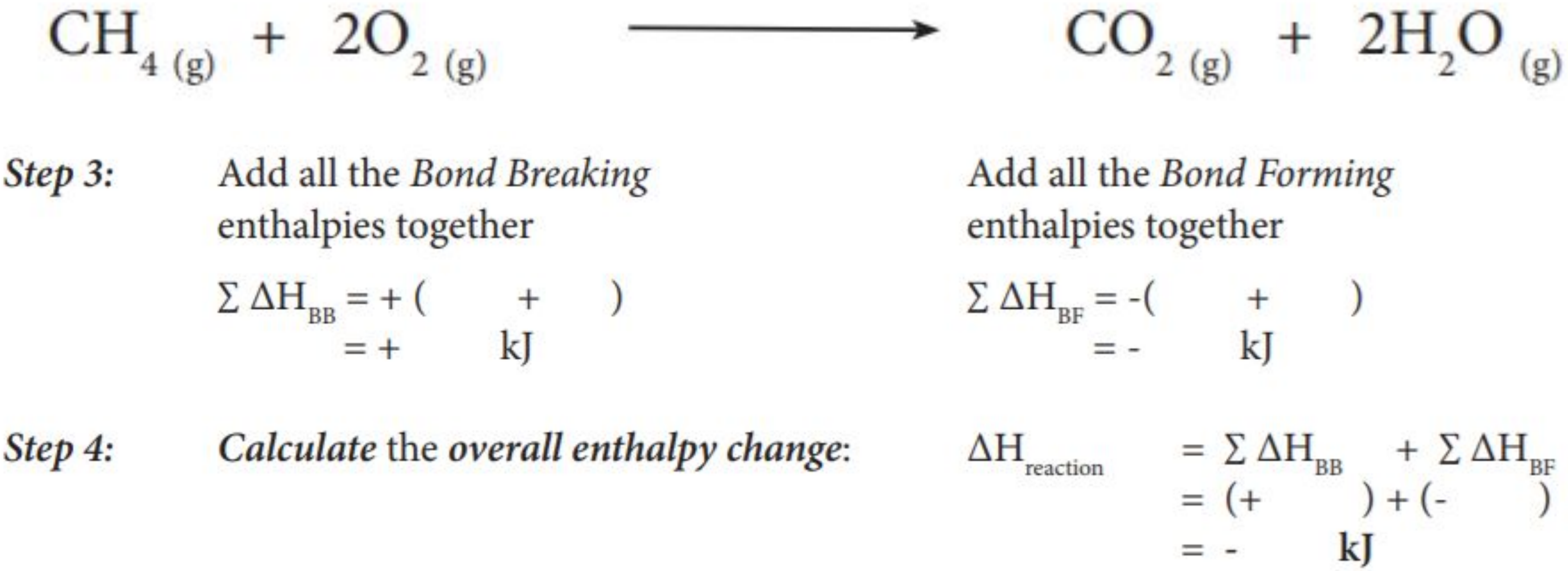
x C—H = x = + kJ
x O=O = x = + kJ

List the ***number*** and ***type of bonds*** being ***formed*** and calculate the ***enthalpies*** (table on *page 40*).

(exothermic)
(negative)

x C=O = x - = - kJ
x O—H = x - = - kJ

WALK THROUGH EXAMPLE



$$CH_{4\,(g)} + 2O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2H_2O_{(g)}$$

Step 3: Add all the *Bond Breaking* enthalpies together

$\Sigma\,\Delta H_{BB} = +(\quad\quad + \quad\quad)$

$= + \quad\quad \text{kJ}$

Add all the *Bond Forming* enthalpies together

$\Sigma\,\Delta H_{BF} = -(\quad\quad + \quad\quad)$

$= - \quad\quad \text{kJ}$

Step 4: ***Calculate*** the ***overall enthalpy change***:

$\Delta H_{reaction} = \Sigma\,\Delta H_{BB} + \Sigma\,\Delta H_{BF}$

$= (+\quad\quad) + (-\quad\quad)$

$= - \quad\quad$ **kJ**

MNEMONIC

You can remember how to calculate the overall reaction energy using the mnemonic:

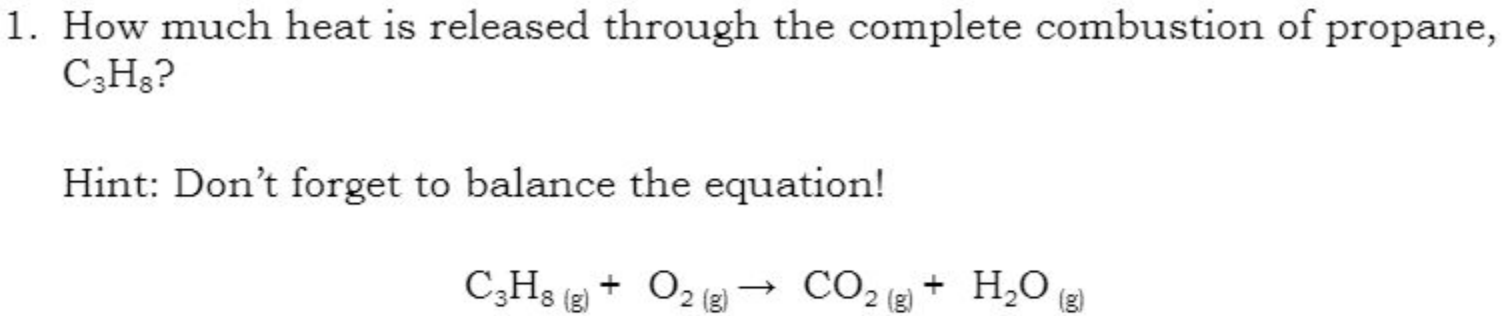
Basket **B**all is **M**y **B**est **F**riend

Bonds Broken Minus Bonds Formed

$$\Delta H_{\text{reaction}} = \underset{\substack{\text{(endothermic)} \\ \text{(positive)}}}{\sum \Delta H_{BB}} + \underset{\substack{\text{(exothermic)} \\ \text{(negative)}}}{\sum \Delta H_{BF}}$$

Videos

TOGETHER (BOND ENTHALPY TABLE ON PAGE 1 OF PRACTICE PROBLEMS)

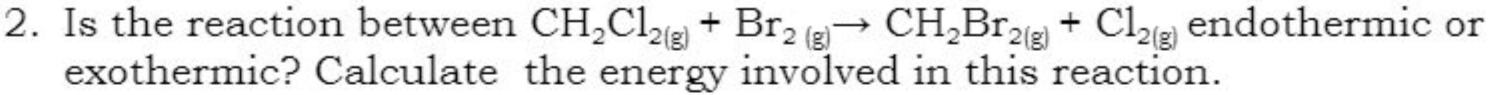


1. How much heat is released through the complete combustion of propane, C_3H_8?

   Hint: Don't forget to balance the equation!

$$C_3H_{8\,(g)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{(g)}$$

YOu Do, We Review



2. Is the reaction between $CH_2Cl_{2(g)}$ + $Br_{2\ (g)} \rightarrow$ $CH_2Br_{2(g)}$ + $Cl_{2(g)}$ endothermic or exothermic? Calculate the energy involved in this reaction.

6.8 Enthalpy of Formation

ENDURING UNDERSTANDING:	
ENE-3	The energy exchanged in a chemical transformation is required to break and form bonds.
LEARNING OBJECTIVE:	
ENE-3.B	Calculate the enthalpy change for a chemical or physical process based on the standard enthalpies of formation.

Bond Breaking and Formation

In order for a chemical reaction to occur bonds need to be broken and bonds need to be formed.

Bond breaking

- endothermic process
- ΔH is +
- energy is transferred from the surroundings to the system

Bond forming

- exothermic process
- ΔH is -
- energy is transferred from the system to the surroundings

Standard Enthalpy

Standard enthalpy is abbreviated as ΔH°.

The ° tells us that we are dealing with the **standard state, which is 1 mol, 1 atm, and typically 298 K (25°C** which is approximately room temperature).

In 6.7 we learned how to use the bond energies to calculate the enthalpy for a reaction; this section allows us to find the enthalpy for a reaction by looking at the **energy of a molecule as a whole** entity.

Subscripts

The enthalpy of formation at standard state, $\Delta H°_f$, is the energy change that occurs with the formation of 1 mole of a substance.

The enthalpy of reaction at standard state, ΔH°rxn is the energy change that occurs for a reaction.

Different subscripts like “rxn” or “f” indicate the type of enthalpy we are measuring

- **Anytime you see ° you need to verify you are working with the elements in their standard states!**

STANDARD STATES FILL IN THE BLANKS

Enthalpy of Formation: is the *energy change* when ***1 mole*** of a ***substance is formed from its constituent elements*** in their ***standard states***

$$+ \qquad + \qquad \longrightarrow \qquad H_2SO_{4\,(l)} \qquad \Delta H^\circ_{form} = -811\ \text{kJ mol}^{-1}$$

$$+ \qquad \longrightarrow \qquad HI_{(g)} \qquad \Delta H^\circ_{form} = +26.5\ \text{kJ mol}^{-1}$$

☞ by definition, the enthalpy of formation of ***an element*** = 0 kJ mol^{-1}

☞ a ***negative*** ΔH°_{form} suggests substance is ***more stable*** than its elements

☞ a ***positive*** ΔH°_{form} suggests substance is ***less stable*** than its elements

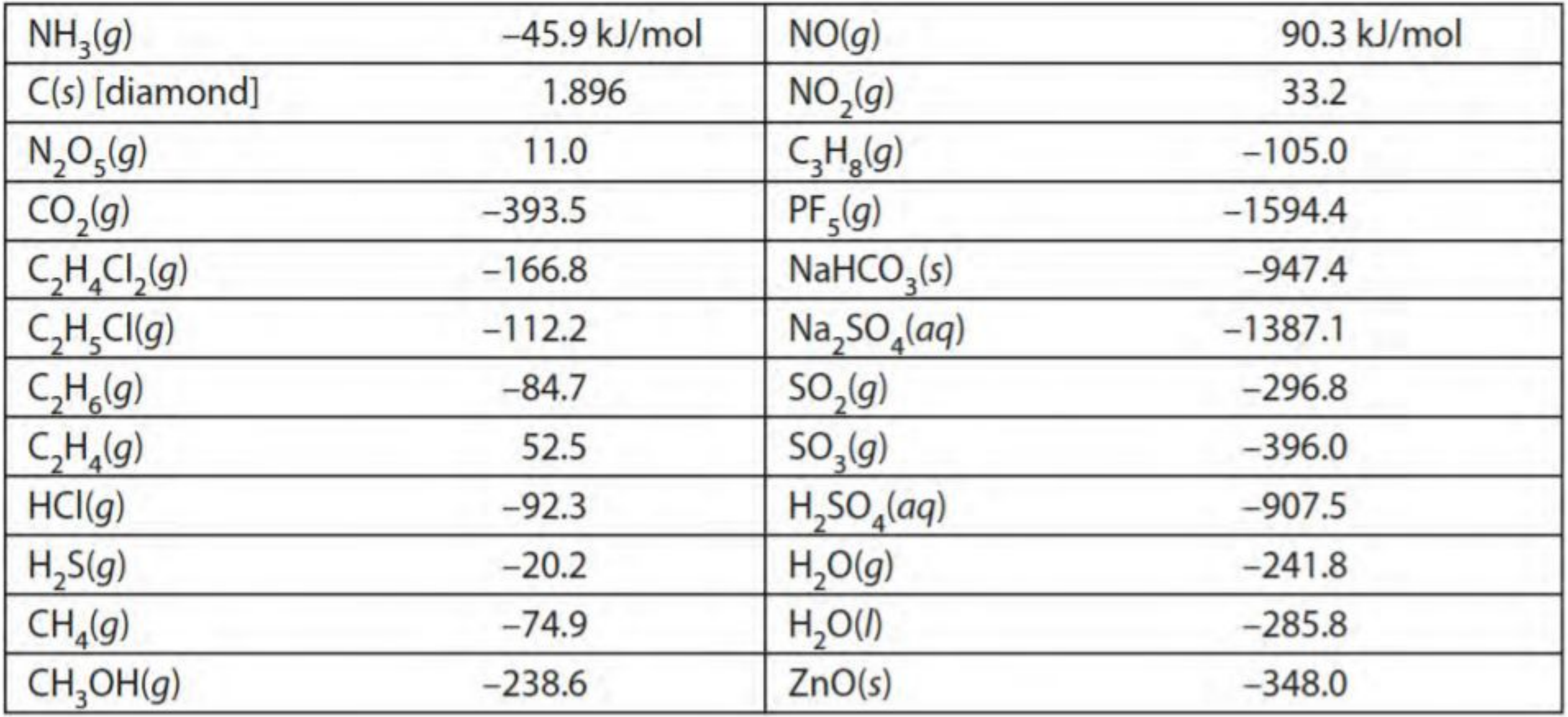
Same Equation (on Equation Sheet)

$$\Delta H^\circ_{reaction} = \Sigma\Delta H_f{}^\circ{}_{products} - \Sigma\Delta H_f{}^\circ{}_{reactants}$$

Enthalpy of a reaction is the equal to the sum of the formation enthalpies of the products minus the sum of the formation enthalpies of the reactants.

- **Be sure to account for the moles in the balanced equation.**

ENTHALPIES OF FORMATION



$NH_3(g)$	–45.9 kJ/mol	$NO(g)$	90.3 kJ/mol
$C(s)$ [diamond]	1.896	$NO_2(g)$	33.2
$N_2O_5(g)$	11.0	$C_3H_8(g)$	–105.0
$CO_2(g)$	–393.5	$PF_5(g)$	–1594.4
$C_2H_4Cl_2(g)$	–166.8	$NaHCO_3(s)$	–947.4
$C_2H_5Cl(g)$	–112.2	$Na_2SO_4(aq)$	–1387.1
$C_2H_6(g)$	–84.7	$SO_2(g)$	–296.8
$C_2H_4(g)$	52.5	$SO_3(g)$	–396.0
$HCl(g)$	–92.3	$H_2SO_4(aq)$	–907.5
$H_2S(g)$	–20.2	$H_2O(g)$	–241.8
$CH_4(g)$	–74.9	$H_2O(l)$	–285.8
$CH_3OH(g)$	–238.6	$ZnO(s)$	–348.0

EXAMPLE

$$\Delta H_{reaction} = \sum \Delta H^{\circ}_{f\ products} - \sum \Delta H^{\circ}_{f\ reactants}$$

For example,

$$CH_{4\,(g)} + 2O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2H_2O_{(g)}$$

Step 1: Add up all the enthalpies of formation of the ***reactants***.

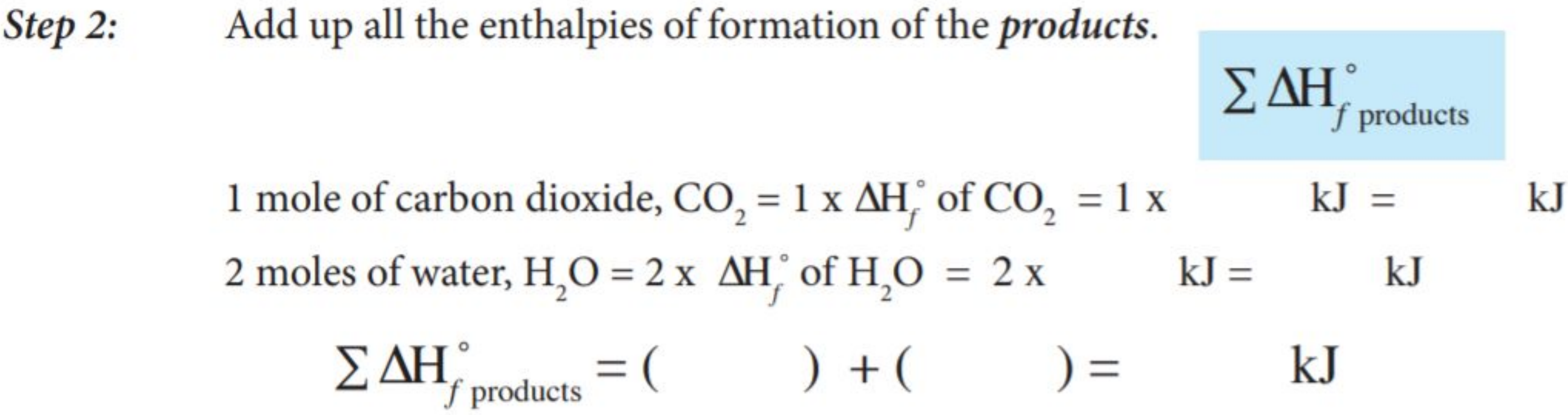
$$\sum \Delta H^{\circ}_{f\ reactants}$$

1 mole of methane, $CH_4 = 1 \times \Delta H^{\circ}_f$ of $CH_4 = 1 \times$ kJ = kJ

2 moles of oxygen, $O_2 = 2 \times \Delta H^{\circ}_f$ of $O_2 = 2 \times$ kJ = kJ

$$\sum \Delta H^{\circ}_{f\ reactants} = (\qquad) + \quad = \qquad kJ$$

EXAMPLE



Step 2: Add up all the enthalpies of formation of the ***products***.

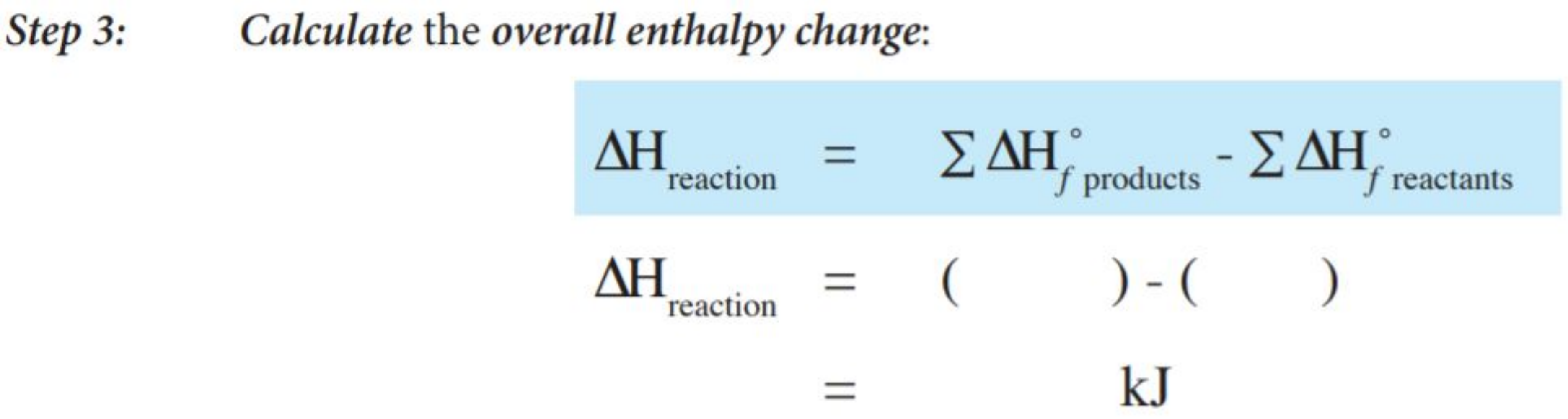
$\sum \Delta H^\circ_{f\ \text{products}}$

1 mole of carbon dioxide, $CO_2 = 1 \times \Delta H^\circ_f$ of $CO_2 = 1 \times \quad$ kJ $= \quad$ kJ

2 moles of water, $H_2O = 2 \times \Delta H^\circ_f$ of $H_2O = 2 \times \quad$ kJ $= \quad$ kJ

$$\sum \Delta H^\circ_{f\ \text{products}} = (\qquad) + (\qquad) = \qquad \text{kJ}$$

Example



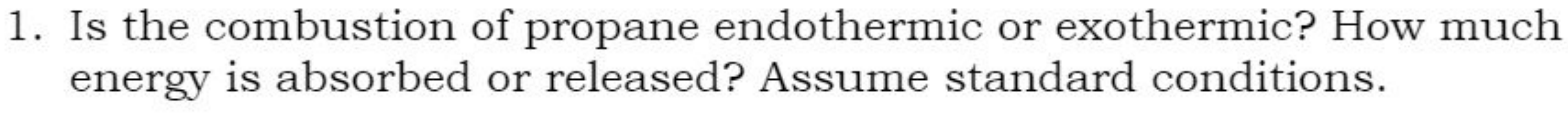
Step 3: ***Calculate*** the ***overall enthalpy change***:

$$\Delta H_{\text{reaction}} = \sum \Delta H^{\circ}_{f\ \text{products}} - \sum \Delta H^{\circ}_{f\ \text{reactants}}$$

$$\Delta H_{\text{reaction}} = (\qquad\qquad) - (\qquad\qquad)$$

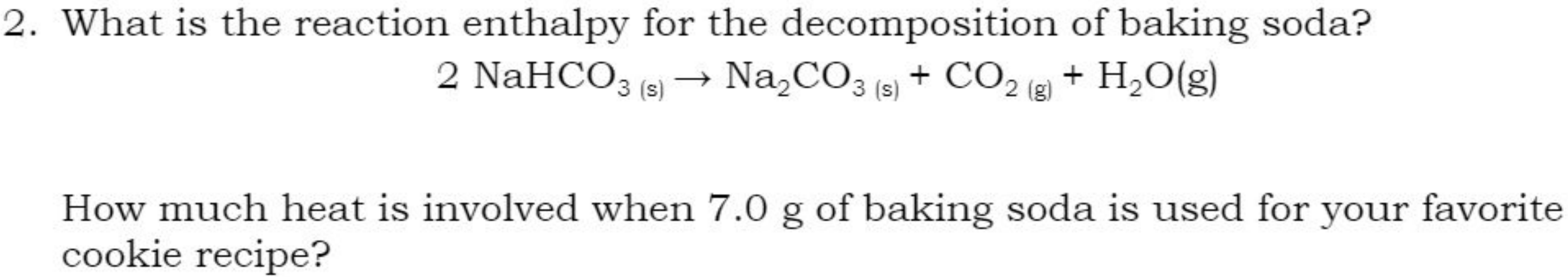
$$= \qquad\qquad \text{kJ}$$

TOGETHER



1. Is the combustion of propane endothermic or exothermic? How much energy is absorbed or released? Assume standard conditions.

YOU DO, WE REVIEW



2. What is the reaction enthalpy for the decomposition of baking soda?

$$2\ NaHCO_{3\ (s)} \rightarrow Na_2CO_{3\ (s)} + CO_{2\ (g)} + H_2O(g)$$

How much heat is involved when 7.0 g of baking soda is used for your favorite cookie recipe?

6.9 Hess's Law

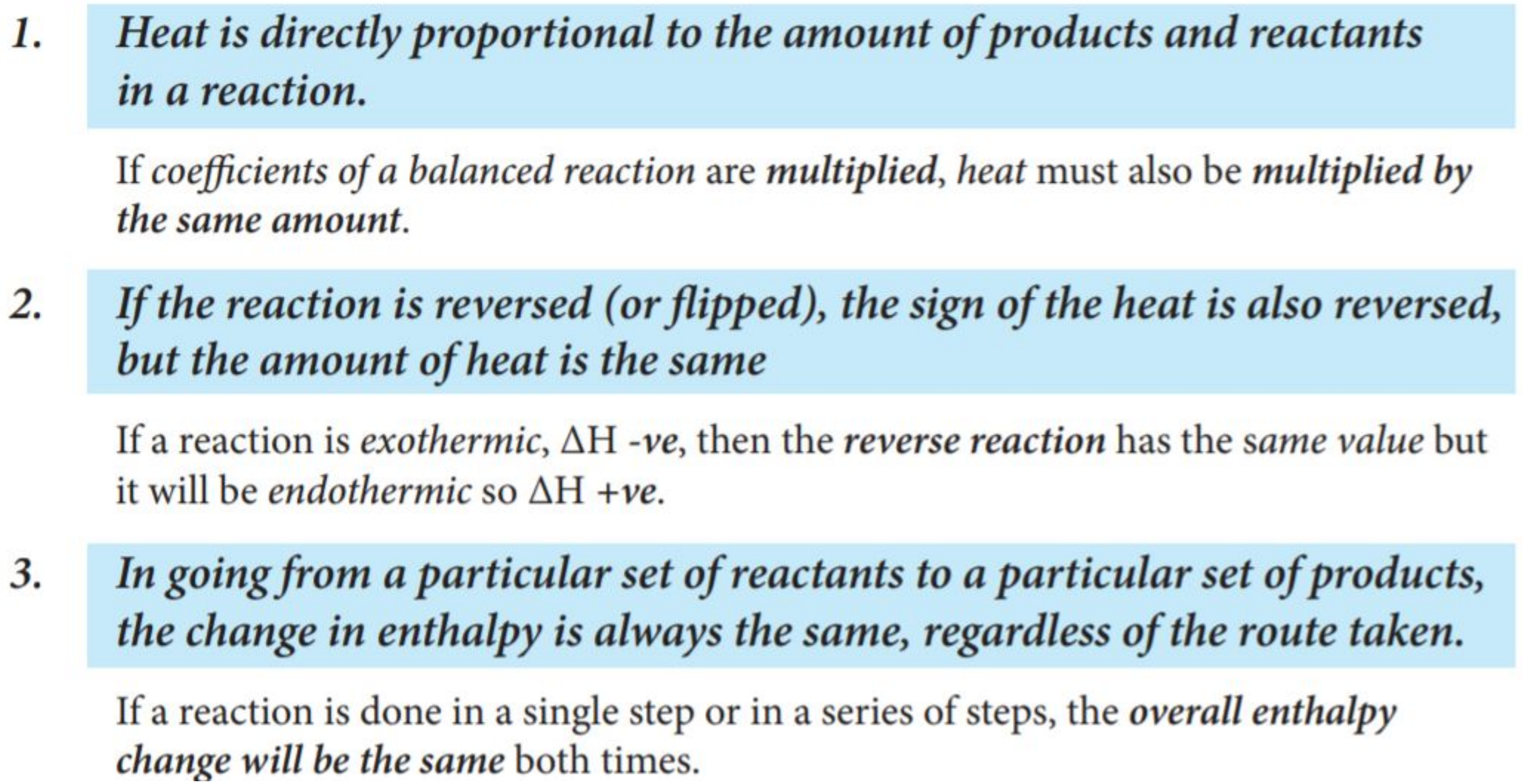
ENDURING UNDERSTANDING:

ENE-3	The energy exchanged in a chemical transformation is required to break and form bonds.

LEARNING OBJECTIVE:

ENE-3.C	Represent a chemical or physical process as a sequence of steps.

HESS'S LAW: 3 PROPERTIES



1. ***Heat is directly proportional to the amount of products and reactants in a reaction.***

   If *coefficients of a balanced reaction* are ***multiplied***, *heat* must also be ***multiplied by the same amount***.

2. ***If the reaction is reversed (or flipped), the sign of the heat is also reversed, but the amount of heat is the same***

   If a reaction is *exothermic*, ΔH -*ve*, then the ***reverse reaction*** has the *same value* but it will be *endothermic* so ΔH +*ve*.

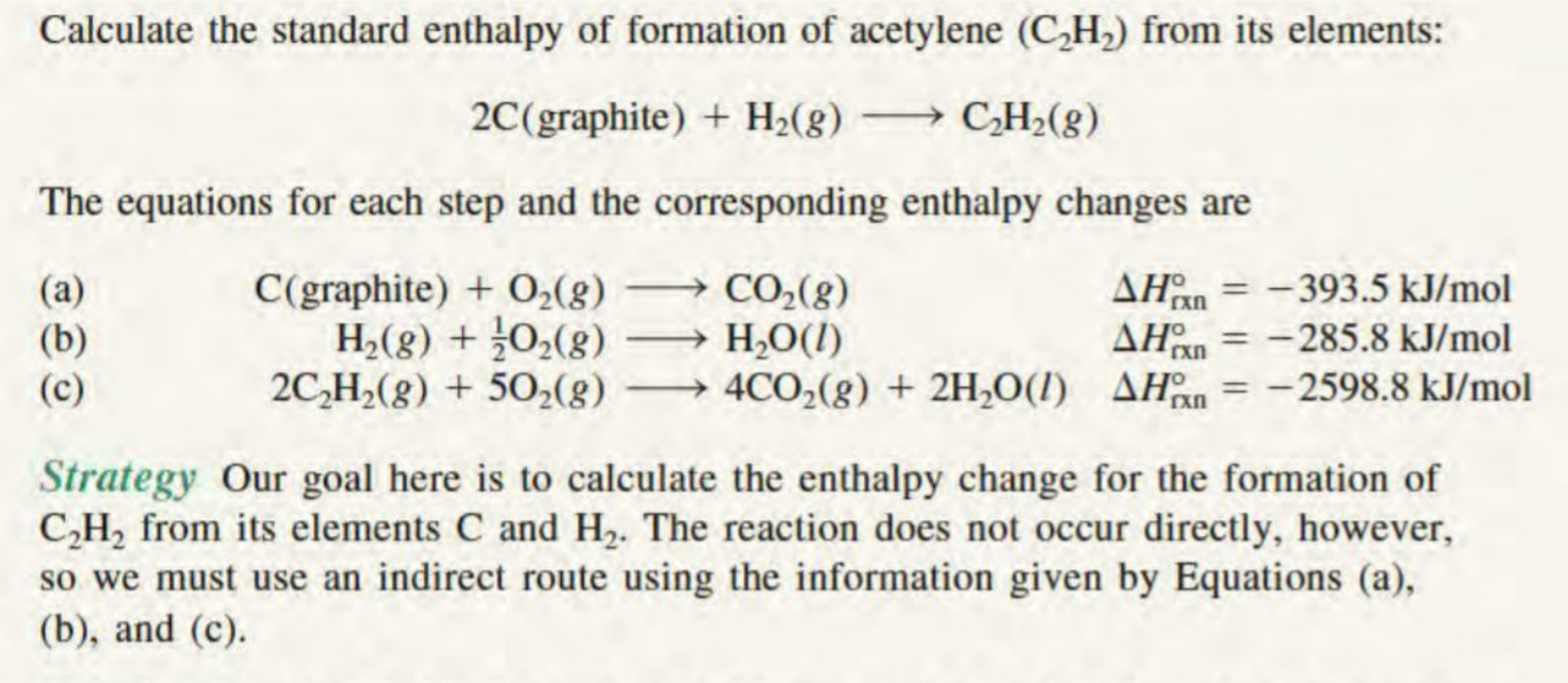
3. ***In going from a particular set of reactants to a particular set of products, the change in enthalpy is always the same, regardless of the route taken.***

   If a reaction is done in a single step or in a series of steps, the ***overall enthalpy change will be the same*** both times.

#3 DESCRIBES STATE FUNCTIONS

Hess’s Law is based on the idea that enthalpy is a state function, state functions are beyond the scope of AP chemistry, but the idea of a state function essentially means that the **numerical value of a property doesn’t depend on the path taken to find that value.**

EASIEST TO WORK THROUGH AN EXAMPLE!



Calculate the standard enthalpy of formation of acetylene (C_2H_2) from its elements:

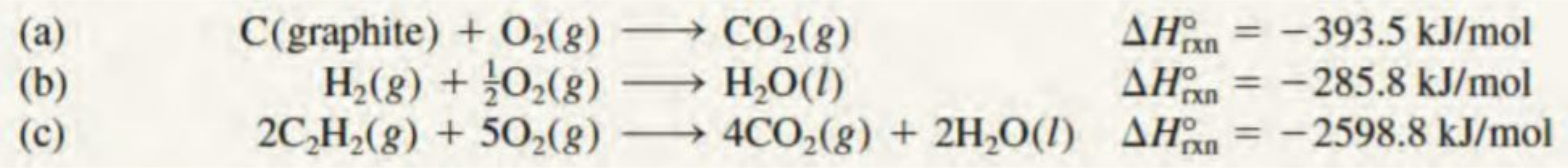
$$2C(\text{graphite}) + H_2(g) \longrightarrow C_2H_2(g)$$

The equations for each step and the corresponding enthalpy changes are

$$\begin{array}{llll}
\text{(a)} & C(\text{graphite}) + O_2(g) \longrightarrow CO_2(g) & & \Delta H^\circ_{\text{rxn}} = -393.5 \text{ kJ/mol} \\
\text{(b)} & H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l) & & \Delta H^\circ_{\text{rxn}} = -285.8 \text{ kJ/mol} \\
\text{(c)} & 2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(l) & & \Delta H^\circ_{\text{rxn}} = -2598.8 \text{ kJ/mol}
\end{array}$$

Strategy Our goal here is to calculate the enthalpy change for the formation of C_2H_2 from its elements C and H_2. The reaction does not occur directly, however, so we must use an indirect route using the information given by Equations (a), (b), and (c).

$2C\ (\mathrm{GRAPHITE}) + H_{2(G)} \longrightarrow C_2H_{2(G)}$



(a) $C(\text{graphite}) + O_2(g) \longrightarrow CO_2(g)$ $\Delta H^\circ_{\text{rxn}} = -393.5\ \text{kJ/mol}$

(b) $H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l)$ $\Delta H^\circ_{\text{rxn}} = -285.8\ \text{kJ/mol}$

(c) $2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(l)$ $\Delta H^\circ_{\text{rxn}} = -2598.8\ \text{kJ/mol}$

1. How many moles of graphite are needed to produce ethyne (IUPAC name for the common name acetylene)?
2. How many moles of Hydrogen gas?
3. How many moles of ethyne are produced? AND we need to reverse the reaction!

WHAT DOES THE RESULT MEAN?

The ΔH_f ethyne = ΔH_{rxn} = 226.6 kJ/mol

When 1 mole of C_2H_2 is formed from 2 moles of C and 1 mole of H_2, 226.6 kJ of heat are absorbed from the surroundings.

This is an endothermic reaction!

TOGETHER

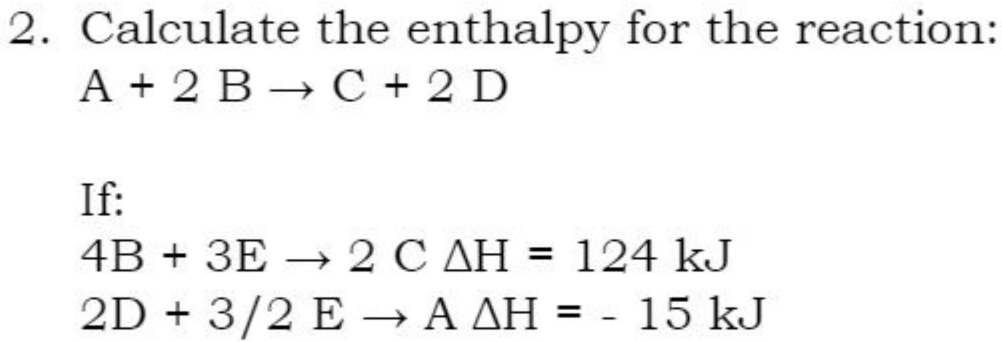
1. Given the following reactions:

   $2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$ $\Delta H = -116$ kJ
   $2N_{2(g)} + 5O_{2(g)} + 2H_2O_{(l)} \rightarrow 4HNO_{3(aq)}$ $\Delta H = -256$ kJ
   $N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$ $\Delta H = +183$ kJ

   Calculate the enthalpy change for the reaction below:

   $3NO_{2(g)} + H_2O_{(l)} \rightarrow 2HNO_{3(aq)} + NO_{(g)}$ $\Delta H = ???$

YOu Do, We Review



2. Calculate the enthalpy for the reaction:
   $A + 2\,B \rightarrow C + 2\,D$

   If:
   $4B + 3E \rightarrow 2\,C$ ΔH = 124 kJ
   $2D + 3/2\,E \rightarrow A$ ΔH = - 15 kJ