Thermodynamics: Chemistry 30 Diploma Preparation

Two major outcomes of this unit:

- Determine and interpret **energy changes** in chemical reactions
- Explain and communicate **energy changes** in chemical reactions

Fourteen major points you MUST know.

1. Q = mcΔt (shows the analysis of heat transfer) This is a kinetic change
2. Explain how stored energy in chemical bonds of hydrocarbons originated at the sun.
3. Define enthalpy (kJ) [ΔH] and molar enthalpy (kJ/mol) [H] for chemical reactions
4. Write balanced chemical reactions that include energy changes
5. Use and interpret ΔH to communicate and calculate energy changes in chemical reactions
6. Use and explain Hess's law to calculate energy changes for a net reaction from a series of reactions.
7. Use Calorimetry data to determine the enthalpy changes in chemical reactions.
8. Identify that $H_2O_{(l)}$ and $CO_{2(g)}$ are the reactants in photosynthesis and the products of cellular respiration.
9. Identify that $H_2O_{(g)}$ and $CO_{2(g)}$ are the products of hydrocarbon combustion in an open system.
10. Classify chemical reactions as endothermic or exothermic
11. Define activation energy as the energy barrier that must be overcome for a chemical reaction to occur.
12. Explain energy changes that occur during chemical reactions, referring to bonds breaking and forming as well as by changes in potential and kinetic energy.
13. Analyze and label energy diagrams of chemical reactions, including reactants, products, enthalpy changes and activation energy.
14. Explain that catalysts increase the reaction rates by providing alternative pathways for changes without affecting the net amount of energy involved.

Thermo chemistry- The study of energy changes by a chemical system during a chemical reaction (This can include both kinetic changes and potential changes)

- Enthalpy is a value that reflects both the kinetic and potential energy of a system
- Enthalpy is measured using a symbol "H"
- $H = E_k + E_p$

Comparison of kinetic energy and potential energy

Kinetic energy KE (movement)

- Sometimes referred to a "Q"
- Energy of motion
- Electrons moving within atoms
- Vibration of atoms connected by chemical bonds
- Rotation and translation of molecules
- We will measure this by observing a change of temperature of the system.

Potential energy PE (Position)

- Bonds are breaking and reforming.
- Intramolecular bonds – forces within atoms or molecules. Covalent and ionic bonds
- Intermolecular bonds – forces between different atoms or molecules
- We want to find the energy required or released when the bonding changes.

ΔH is the change of enthalpy
This measures the difference between the potential & chemical energy at the end and the potential & kinetic energy at the beginning of a reaction.

Simple application of thermodynamics:

I. Let's consider situations where the chemical does not change its potential. The chemical will change its kinetic energy only.
- This will require a change of temperature while the chemical remains in the same state and does not change its bonding structure.
- We will use a formula that focuses JUST on kinetic energy.
- Sometimes the symbol Q is used to represent this kinetic energy.

  - $Q = mc\Delta t$
  - Q = total kinetic energy
  - Q can be negative (heat energy is removed …. the substance is going through an EXOTHERMIC change)
  - Q can be positive (heat energy is added …. The substance is going through an ENDOTHERMIC change)
  - m = mass of substance (mg, g, kg, Mg, etc)
  - c = specific heat capacity of substance $\frac{J}{g.°C}$ or $\frac{kJ}{kg°C}$
    - see page 3 of data booklet for a smattering of specific heat capacities

- These values are <u>defined</u> for pure substances

- Δt = final temperature – initial temperature
  - Δt can be negative if the temperature drops (exothermic)
  - Δt can be positive if the temperature rises (endothermic)

A similar spin off of this formula is
$Q = C\Delta t$

Here C = heat capacity in $\frac{J}{°C}$ or $\frac{kJ}{°C}$. This value will be valid for a mixture of items that all change temperature together. (Like a calorimeter)

Sample questions with worked solutions:

1. A 25.7 g sample of aluminium has energy added to change its temperature from 14.6 °C to 59.1 °C. Find the change of enthalpy.

   Note that aluminium has not changed its potential energy (state) between these two temperatures, so the ΔH is only a kinetic change.

   $Q = mc\Delta t$

   $Q = 25.7\ g \times 0.897\ \frac{J}{g\cdot°C} \times (59.1\ °C - 14.6\ °C)$

   Q = 1025.85…. J or 1.03 kJ of energy is added

2. Suppose that 750 J of energy is added to a 50.0 g sample of water at 8.20 °C. What will be the final temperature of the water sample?

   Note that water will not change its potential when such a small amount of energy is added to this cold sample. So again

   $Q = mc\Delta t$ $\qquad\qquad$ $\Delta t = t_f - t_i$

   $\Delta t = \frac{Q}{mc}$ $\qquad\qquad$ $t_f = \Delta t + t_i$

   $\Delta t = \frac{750J}{50.0g x 4.19 \frac{J}{g°C}}$ $\qquad\qquad$ $t_f = 3.5799\ °C + 8.20\ °C$

   $\Delta t = 3.5799....\ °C$ $\qquad$ $t_f = 11.8\ °C$ is the final temperature

3. What mass of water will require 80 kJ of heat to warm from 12 °C to 30 °C?

   $m = \frac{Q}{c\Delta t}$

m = $\dfrac{80kJ}{(30°C - 12°C)x4.19\dfrac{kJ}{kg°C}}$

m = 1.0607 kg

m = 1.1 kg of water will be heated

4. A mixture of liquids in a container has a heat capacity of 4.31 $\dfrac{kJ}{°C}$. How much heat will be needed to heat this mixture 45.0°C?

$Q = C \Delta t$

$Q = 4.31 \dfrac{kJ}{°C} . X 45.0°C$

Q = 193.95 kJ

Q = 194 kJ

Independent Practice questions:

1. Find the Δt if a 56 g sample of water gains 1.2 kJ of energy. (5.1°C)

2. A metal sample takes in 590 J of energy. If the sample has a mass of 25.0g and goes from an initial temperature of 12.0°C to a final temperature of 82.1°C, find the specific heat capacity in $\dfrac{J}{g.°C}$ (0.337 $\dfrac{J}{g.°C}$)

3. A 55.0 g sample of water at 75.4°C has 950 J of energy removed from it. Find the final temperature of the water sample. (71.3°C)

4. When a mixture of liquids in a container is cooled by 20.0°C, 74.6 kJ of energy is removed. Find the heat capacity in $\dfrac{kJ}{°C}$. (3.73 $\dfrac{kJ}{°C}$)

Note that up until now we have only talked about one substance or a mixture of things that has changed its kinetic energy only!

II. Sometimes we need to look at two substances.
They could both be going through a temperature change only (kinetic only) OR one could be going through a potential change while the other goes through a kinetic change.

A process that lets us investigate one substance by looking at another substance is called Calorimetry.

- A calorimeter is an isolated system.
- Calorimeter is an instrument that allows a change to occur to one chemical (either kinetic or potential) and the energy change for this chemical is reflected by a volume of surrounding water.
- Simple calorimeters are nested Styrofoam cups.
- Expensive calorimeters are called "bomb" calorimeters. These have a reaction chamber that is surrounded by a volume of water.

Assumptions that can be made for a good calorimeter

- Any energy gained or lost by the reaction is reflected by a gain or loss to the fluid in the calorimeter.
- Matter is conserved in the reaction.
- Any dilute solutions are treated as if they are water and will also have a "c " value of 4.19 $\frac{J}{g.°C}$.
- Density of water and all dilute solutions will be considered to be identical at 1.00 $\frac{g}{mL}$
- The amount of heat gained or lost by the calorimeter itself is thought to be negligible. This does not refer to the liquid contents of the calorimeter, but to the actual instrument itself.
- The amount of energy transferred in a reaction is measured as **heat** (Joules, kilojoules, etc)
    - If combustion reactions are done in a calorimeter, the water is assumed to come off as $H_2O_{(l)}$ because of the low temperatures.
    - If combustion reactions are done in an open setting like a barbeque, then water will be assumed to come off as $H_2O_{(g)}$.

Note that when using a calorimeter, some chemical(s) is/are undergoing a POTENTIAL change OR a KINETIC change while the 'water' of the calorimeter is reflecting the energy released or required by going through a change of temperature

Since we are now looking at TWO substances, the process for solving the problems will be a bit more complicated.

The laws of thermodynamics state that in a closed system energy cannot be created or destroyed. The energy is in fact conserved.

So Energy gained by one part of the system must have been lost by another part of the system.

1. When a reaction takes in heat, the water in the calorimeter will cool. So for the reaction, ΔH_r will be a positive value (endothermic)
   The water will lose heat Q, so the value for $Q = mc\Delta t$ will be negative (exothermic)

OR

2. When a reaction loses heat, the water in the calorimeter will warm up . So for the reaction, ΔH_r will be a negative value (exothermic)The water will gain heat Q, so the value for $Q = mc\Delta t$ will be positive (endothermic)

According to the law of thermodynamics, heat lost and heat gained are equal. The sum of identical positive and negative values will always be zero.

	Reaction + calorimeter = 0
	$\Delta H_r + Q = 0$
Chemical has a potential change	$nH + mc\Delta t = 0$
Chemical has a kinetic change	$mc\Delta t + mc\Delta t = 0$

Example 1:

Sufficient hydrogen and oxygen gas are introduced into the reaction chamber of a bomb calorimeter to produce 5.00 g of water.

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

Chemical is going through a POTENTIAL change

Surrounding the calorimeter, the 750 g of water in the calorimeter goes from an initial temperature of 13.4 °C to 38.6 °C. Find the molar heat for the formation of water.

$nH + mc\Delta t = 0$

$$\left[\frac{5.00g}{18.02\frac{g}{mol}} \text{ x H}\right] + \left[0.750 \text{ kg x } 4.19 \frac{kJ}{kg°C} \text{ x } (38.6\ °C - 13.4\ °C)\right] = 0$$

H = -285 kJ/mol ….. This means the reaction was exothermic and the calorimeter gained the heat!

Example 2

A 1.00 g sample of propane undergoes complete combustion to form carbon dioxide and water vapour. If the molar heat of combustion for propane is -2219.9 kJ/mol, what mass of water in the calorimeter will go through a temperature increase of 59.0 °C

$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)}$

Chemical is going through a POTENTIAL change

Note: combustion reactions is done in a calorimeter so the water is assumed to come off as $H_2O_{(l)}$ because of the low temperatures

$$nH + mc\Delta t = 0$$

$$\left[\frac{1.00g}{44.11\frac{g}{mol}} \text{ x } -2219.9 \frac{kJ}{mol}\right] + \left[m \text{ x } 4.19 \frac{kJ}{kg°C} \text{ x } 59.0\ °C\right] = 0$$

m = 0.20357 kg or 204 g of water

Example 3

A sample of water at 15.0°C is mixed with 150 g of water in a calorimeter at 70.5°C. The mixture of the two samples comes to a final equilibrium temperature of 55.9°C. Find the mass of the first sample of water.

Notice Chemical has a kinetic change

$mc\Delta t + mc\Delta t = 0$

$$[m \times 4.19 \frac{kJ}{kg^{\circ}C} \times (55.9^{\circ}C - 15.0^{\circ}C)] + [150\ g \times 4.19 \frac{kJ}{kg^{\circ}C} \times (55.9^{\circ}C - 70.5^{\circ}C)] = 0$$

m = 53.5 g of water.

Independent practice.

Rules:

1. Write and balance a reaction.
   - Do the substances change chemically and take up a new position?
     - o $nH + mc\Delta t = 0$
   - Do the substances just change temperature and keep their old position?
     - o $mc\Delta t + mc\Delta t = 0$
2. Establish if the calorimeter is the endothermic or exothermic step. (Does the temperature rise or fall?)
3. The substance in the calorimeter will do the opposite of the calorimeter.
   - If the calorimeter is endothermic, the reaction is exothermic.
   - If the calorimeter is exothermic, the reaction is endothermic.
4. Use the formulas above to help solve for the missing information.

Questions to practice

1. The complete combustion of a 20 g sample of butane heats the water in a calorimeter by 40 °C. If the molar enthalpy of combustion for butane is -2877.3 kJ/mol, find the mass of water in the calorimeter. (5.9 kg)

2. A sample of ethanol is combusted and heats a calorimeter of 6.1 L of water by 10°C. If the molar enthalpy of combustion for ethanol is -1366.8 kJ/mol find the mass of ethanol that is combusted. (8.6 g)

3. In the lab, two chemistry 30 students dissolve 0.150 mol of an ionic compound in 100 mL of water. If the water decreases in temperature by 8.25°C, then find the molar enthalpy of dissolving for the ionic compound is __________ kJ/mol. (+23.0 kJ/mol)

4. A 190 mL sample of water at 45.0°C is mixed with 250 mL of water at 12.0°C. Find the equilibrium temperature of this mixture. (26.3°C)

5. A 100 mL sample of 2.00 mol/L sample of hydrochloric acid neutralizes a sample of solid $Ba(OH)_{2(s)}$. If the molar enthalpy of neutralization for hydrochloric acid is -148.65 kJ/mol and the initial temperature of the solution is 10.4°C, find the final temperature of the solution. (81.4°C)

6. A 50 g sample of octane combusts to heat the water in a calorimeter by 10 °C. If the molar enthalpy of combustion for octane is -5470.1 kJ/mol, find the mass of water in the calorimeter.(57 kg)

7. A 1.70 kg sample of methanol is combusted and heats a calorimeter of 0.500 kL of water by 18.0°C. Find the molar enthalpy of combustion of methanol. (-711 kJ/mol)

8. In the lab, two chemistry 30 students dissolve 70 g of sodium hydroxide in 500 mL of water. If the water increases in temperature by 51°C, then find the molar enthalpy of dissolving for sodium hydroxide. Express the answer in kJ/mol. (-61 kJ/mol)

9. A 150 mL sample of 0.250 mol/L hydrochloric acid reacts with a sample of solid $K_{(s)}$. If the molar enthalpy of reaction for hydrochloric acid is -344.2 kJ/mol and the final temperature of the solution is 81.9°C, find the initial temperature of the solution.
(61.4°C)

10. The molar enthalpy of decomposition of sucrose is 2226.1 kJ/mol**.** What mass of sucrose is decomposed if a calorimeter holding 7.00 L of water has its temperature lowered by 13.1°C? (59.1 g)

11. Sufficient chlorine reacts with tin to <u>form</u> 14 g of $SnCl_{4(l)}$. If the mass of water in the calorimeter is 5.7 kg, and the initial temperature is 34°C, what will be the final temperature of the water in the calorimeter? Remember to look up the molar heat of formation in your data booklet. (36°C**)**

12. A 900 g sample of water at 25.0°C is mixed with another sample of water at 75.0°C. If the equilibrium temperature of the mixture is 38.0°C, find the mass of the second sample of water. (316 g)

Up until now all the values have been found experimentally in the lab.
This is fun, but very time consuming!!
Fortunately we can use our data book to predict enthalpy of reactions without stepping inside the lab.

Standard Molar Enthalpies of Formation at 298.15 K

These values are valid if we form compounds from their elements at 'standard' conditions.'

- See data booklet pages 5 & 6
- Energy may be required to form compounds from their elements. (benzene +49.1 kJ/mol**,** dinitrogen tetroxide +11.1 kJ/mol, etc)
- Energy may be released if compounds are formed from their elements. (iron (II)oxide -824.2 kJ/mol, zinc sulphide -206.0 kJ/mol)
- The molar enthalpy numbers can be positive. This means energy is absorbed ➜ an endothermic reaction. In an energy diagram, the reactants are below the products
- The molar enthalpy numbers can be negative. This means energy is released ➜ an exothermic reaction. In an energy diagram, the reactants are above the products
- Remember that if a calorimeter is being used, the calorimeter will do the OPPOSITE of the reaction.

Stoichiometry and Formation reactions:

1. Write a balanced reaction for ONE mole of the compound
2. Find the enthalpy of formation in your data booklet.
   a. Remember to write the number on the product side if heat of formation is negative
   b. Remember to write the number on the product side if heat of formation is positive
3. Apply the skills of stoichiometry to solve the question.

EXAMPLE 1:

What mass of potassium chlorate is formed from its elements when 200 kJ of energy is released?

- Note that the enthalpy of formation is NEGATIVE ➜ so put this energy on the product side as it is being RELEASED by the reaction

$$K_{(s)} + \tfrac{1}{2}Cl_{2(g)} + \frac{3}{2}O_{2(g)} \rightarrow \underline{KClO_{3(s)}} + \underline{397.7\text{ kJ}}$$

? g

$$\frac{-397.7kJ}{1mol} = \frac{-200kJ}{?mol}$$

? = 0.5028…. mol

Now moles to mass

m = nM

m = 0.5028 mol x 122.55 g/mol

m = 61.6 g of potassium chlorate

1. How much heat is required to form 150 g of nitrogen monoxide from its elements? (+ 456 kJ)
2. How much heat is released when 100 g of magnesium sulfate is formed from its elements? (- 1.07 x 10^3 kJ)

3. What mass of ammonia is formed from its elements when 500 kJ of energy is released? (186 g)
4. How many moles of zinc oxide is formed from its elements when 59.2 kJ of energy is released? (0.169 mol)
5. How many moles of nitrogen dioxide is formed from its elements when 75 kJ of energy is added? (2.3 mol)
6. When 100 g of silver bromide is formed from its elements, how much energy is released? (-198)

   Life would be good if EVERY reaction was formation …. But of course this is NOT true. So luckily for us a guy called G. H. Hess developed a law to help us with ANY reaction.

1840, *G. H. Hess*, after many experiments, announces that the heat from a chemical reaction is independent of the reaction path (*Hess's* Law)

Theory

Calorimetry can't be used to measure the enthalpy of certain reactions. These reactions generally fall into two categories:

1. **Reaction that can't be isolated from others and occur at the same time.**
   For Example, the formation of carbon monoxide from carbon and oxygen. When carbon burns in air, it forms both carbon dioxide and carbon monoxide. The formation of carbon monoxide can't be separated from that of carbon dioxide. So how can we measure the heat of formation of carbon monoxide?
2. **Reactions that proceed so slowly to effectively measure with a calorimeter.**
   For example, the rusting of iron is a very slow process, and the temperature changes in the surroundings are too small to measure.

In order to measure the heats of these reactions (enthalpies) chemists have developed a theoretical method. The method is based on a simple idea.

Hess's law can be used in several different ways to find the Enthalpy of a reaction. Only the first two ways are required for chemistry 30.

Process 1	Calculate by adding up a series of reactions that will reduce to the required reaction. This is useful when some compounds are NOT in the data booklet!!
Process 2	Calculate using enthalpy of formation data given in the data booklet. Use the formula for Hess's law $\Delta_r H^\circ = \sum [n\Delta_{f\,P}\, H_m^\circ] - \sum [n\Delta_{f\,R}\, H_m^\circ]$
Process 3	Calculating with Hess's Law using Bond Energies (not part of chem. 30)

Process 1: adding up a series of reactions

1) Write the overall reaction that you are attempting to solve for.
2) manipulate the **given equations** until they add up to the overall equation.
3) Combine all the enthalpy information from the given equations and reduce to one value.

4) This value will be the enthalpy for the overall reaction.
5) This is an <u>annoying and cumbersome process</u> …. But you need to be able to do it …. So DIG in.

Examples of process 1

1. Given the following equations:

$4\ NH_{3\,(g)} + 5\ O_{2\,(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{\,(l)}$ $\Delta H° = -1170$ kJ

$4\ NH_{3\,(g)} + 3\ O_{2\,(g)} \rightarrow 2\ N_{2\,(g)} + 6\ H_2O_{\,(l)}$ $\Delta H° = -1530$ kJ

Using these two equations, determine the enthalpy of formation, ($\Delta H°_f$), for nitrogen monoxide, $NO_{(g)}$.
(answer is + 90 kJ)

2. Given the following equations,

- $B_2O_3(s) + 3H_2O(g) \longrightarrow 3O_2(g) + B_2H_6(g)$ $\Delta H = +2035$ kJ
- $H_2O(l) \rightarrow H_2O(g)$ $\Delta H = +44$ kJ
- $H_2(g) + \frac{1}{2}\ O_2(g) \longrightarrow H_2O(l)$ $\Delta H = -286$ kJ
- $2B(s) + 3H_2O_{(l)} \rightarrow \frac{3}{2}\ O_2(g) + B_2H_6\ (g)$ $\Delta H = +\ 126$ kJ

Using these equations, find the enthalpy of formation for $B_2O_3(s)$. (answer is -2041 kJ)

Process 2: Using heats of formation and Hess' formula

$$\boxed{\Delta H = \sum [n\Delta H_f{}^\circ P] - \sum [n\Delta H_f{}^\circ R]}$$

What it is really saying is that the total change of enthalpy (at standard conditions) for a reaction is equal to the SUM of the number of moles *x* the standard enthalpy of formation for all the PRODUCTS – the SUM of the number of moles *x* the standard molar enthalpy of formation for all the REACTANTS.

Practice with Hess's law

1. Use the table of standard enthalpy of formation to find the molar enthalpy for the combustion of methanol.

$$CH_3OH_{(l)} + \frac{3}{2} O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

$$\Delta H = [1 \text{ mol x } -393.5 \text{ kJ/mol} + 2 \text{ mol x } -241.8 \text{ kJ/mol}] - [1 \text{ mol x } -239.2 \text{ kJ/mol} + \frac{3}{2} \text{ mol x } 0 \text{ kJ/mol}]$$

$\Delta H = -637.9$ kJ

If we want to write this into the reaction, it goes on the product side

$$CH_3OH_{(l)} + \frac{3}{2} O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)} + 637.9 \text{ kJ}$$

- Using the information above, find the mass of water that could be heated by 3.7 °C if 3.66 g of methanol is combusted.

  $nH + mc\Delta t = 0$

  $$\left[\frac{3.66g}{32.05\frac{g}{mol}} \text{ x } \frac{-637.9kJ}{mol}\right] + \left[m \text{ x } 4.19 \frac{kJ}{kg^\circ C} \text{ x } 3.7^\circ C\right] = 0$$

  m = 4.7 kg of water can be heated.

2. Use the table of standard enthalpy of formation to write the reaction for the formation of magnesium carbonate from its elements.

- What effect will the formation of 20.0 g of magnesium carbonate have on the temperature of 5.00 L of water? (increase of 12.4°C)

3. Use the table of standard enthalpy of formation to find the molar enthalpy for the combustion of ethane.

- If 0.50 mol of ethane undergoes complete combustion in a calorimeter, what mass of water in a calorimeter will undergo a temperature increase of 17° C? (11 L of water can be heated)

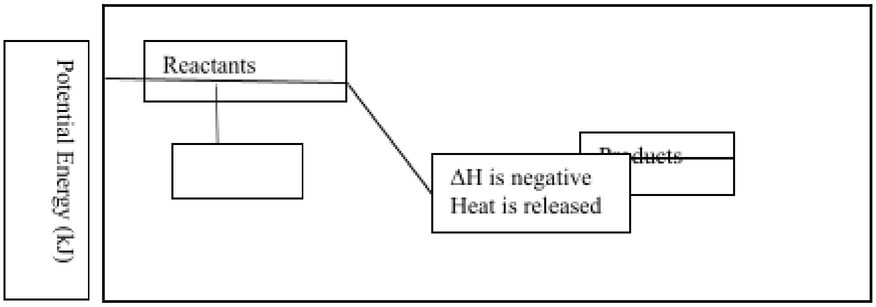
4. Calcium carbonate <u>decomposes</u> into calcium oxide and carbon dioxide in a calorimeter. Find the molar enthalpy for the reaction.

    - What will be the effect on the temperature of 450 mL of water if 2.50 g of calcium carbonate react? (temperature of water will decrease by 17.8°C)

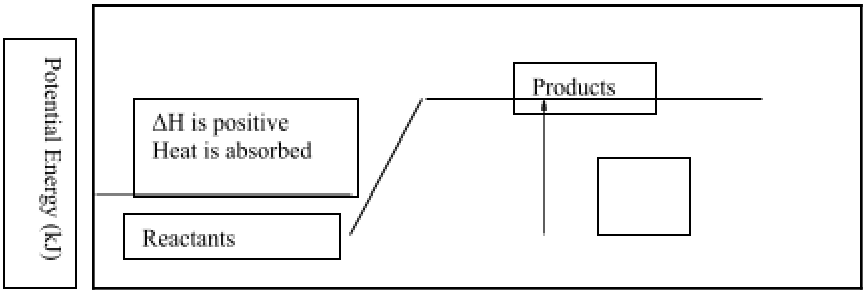
Potential Energy Diagrams:

Simplified diagrams are shown below:

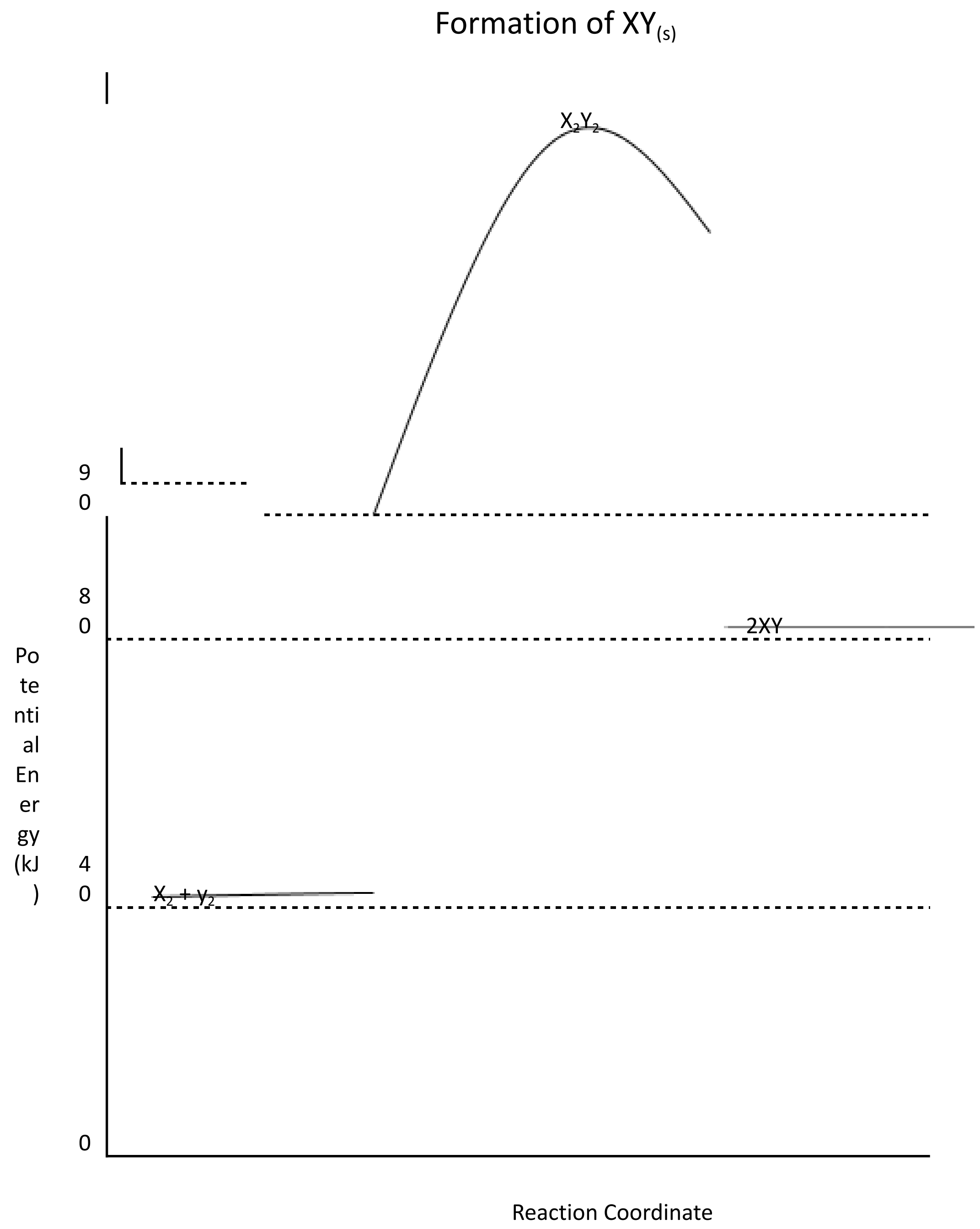
A potential energy graph of an exothermic reaction:



A potential energy graph of an endothermic reaction



But reactions actually have an <u>activation energy</u> that needs to be shown. See the next pages!



The REAL story must show the activation energy.

a) Is the overall reaction as shown exothermic or endothermic? _______________
b) What is the activation energy for the forward reaction? ____________________
c) What is the activation energy for the reverse reaction? ___________________
d) What is the enthalpy change (ΔH) for the forward reaction? ______________
e) What is the enthalpy change (ΔH) for the reverse reaction? _________________
f) Is the reverse reaction exothermic or endothermic? _____________________
g) Which species forms the activated complex? _______________________
h) Which species or set of species has the highest potential energy? ___________________
i) Which species or set of species has the highest kinetic energy? ____________________

j) Which species or set of species has the <u>weakest bonds</u>? ______________
k) Which species or set of species has the <u>strongest bonds</u>? ______________
l) What is the ΔH for the reaction: $X_2Y_2 \rightarrow X_2 + Y_2$? ________________
m) What do you think would be faster, the forward reaction or the reverse reaction? _______________. Give a reason to support your answer. ____
n) Which species or set of species has the <u>lowest kinetic energy</u>? ________
o) Show the ΔH, the activation energy for the forward reaction and the activation energy for the reverse reaction on the graph above.
p) As the reactant particles approach each other before a collision the potential energy goes _____________ while the kinetic energy goes ___________.
q) As the particles of the newly formed products move away from one another, the potential energy goes ________________ while the kinetic energy goes _____________.
r) As reactant molecules approach each other they exert ____________ forces on each other. Thus, as they move together, their speed ____________ and the potential energy ______________.
s) The meaning of "**<u>activation energy"</u>** is _____________________________
____
____

Answers

Assume energy at $x_2 + y_2$ to be 42 kJ
Assume energy at x_2y_2 to be 92 kJ
Assume energy at 2xy to be 82 kJ

a) Endothermic
b) +50 kJ
c) +10 kJ
d) +40 kJ
e) -4- kJ
f) Exothermic
g) x_2y_2
h) x_2y_2
i) $x_2 + y_2$ (they have least potential … therefore most kinetic)
j) x_2y_2 (they have the highest potential … therefore weak bonds)
k) $x_2 + y_2$ (they have low potential, therefore strong bonds)
l) -50 kJ
m) Reverse …. Requires less activation energy to get the reaction started
n) x_2y_2 (low kinetic means high potential energy)
o)
p) Up, down. (as you get closer to a friend, potential for a kiss goes up)
q) Down, up
r) Repulsive, decrease, increase
s) Minimum kinetic energy that reactant molecules must possess in order to have a successful collision and form an activated complex

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Combining Hess's law and energy diagrams.

Write each balanced reaction in three forms:

- **using the Δ H format,**
- **putting the heat into the equation**
- **By sketching a potential energy diagram. Be sure to show a 'bump' for the activation energy.**

a) Complete combustion of glucose
b) single replacement reaction between hydrogen sulfide gas and zinc metal
c) single replacement reaction between silver iodide solid and chlorine gas
d) decomposition of calcium carbonate into calcium oxide solid and carbon dioxide gas
e) addition of oxygen gas to copper (I) oxide solid to make copper (II) oxide

Answers (Hess's law)

a) $C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightarrow 6\ CO_{2(g)} + 6H_2O_{(g)}$ ΔH = -2538.5 kJ

b) $H_2S_{(g)} + Zn_{(s)} \rightarrow H_{2(g)} + ZnS_{(s)}$ ΔH = -185.4 kJ

c) $2AgI_{(s)} + Cl_{2(g)} \rightarrow 2AgCl_{(s)} + I_{2(s)}$ ΔH = -130.4 kJ

d) $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$ ΔH = -146.0 kJ

e) $\frac{1}{2} O_{2(g)} + Cu_2O_{(s)} \rightarrow 2CuO_{(s)}$ ΔH = +179.2 kJ

More Practice Questions:

1. Find the missing values
   - Find the Δt if 500 J of energy is removed from a 25.0 g sample of water.
   - Find the mass of tin that is heated from 12 °C to 74 °C by 3.2×10^2 J of energy.
2. Write a balanced reaction and include the energy to
   - Form 0.5 moles of magnesium sulfate
   - Decompose 2 moles of barium oxide
3. Use the balanced reactions to calculate the amount of energy for each question.
   - $3A + 2B \rightarrow A_3B_2$ ΔH = -24 kJ
     - How many moles of A will be used if 80 kJ of energy is released?
   - $X + YZ \rightarrow Y + XZ$ ΔH = 50 kJ
     - If 2.5 moles of Y is produced, how much energy will be absorbed by the reaction.

4. Find the enthalpy of the reaction

$H_2O_{(g)} + D_2L_{3(s)} \rightarrow 2DH_{(s)} + L_3O_{(s)}$

- Given the reactions below
  1. $2D_2L_{3(s)} + 3O_{2(g)} + 480.3\ kJ \rightarrow 2L_3O_{(g)} + 4DO_{(s)}$
  2. $2DH_{(g)} + 3H_{2(g)} + 3O_{2(g)} + 120.0\ kJ \rightarrow 2DO_{(s)} + 4H_2O_{(g)}$
  3. $H_2O_{(g)} + 241.8\ kJ \rightarrow H_{2(g)} + 0.5\ O_2(g)$

5. Use the Calorimetry data to find the missing values below.
- A 25 g sample of iron reacts with sufficient oxygen to form iron (III) oxide. If the calorimeter around the reaction contains 26 L of water, what will be the Δt for the water in the calorimeter?
- When 200 mL of 1.50 mol/L nitric acid reacts with sufficient lithium hydroxide, the solutions goes from 21.2 °C to 73.8 °C. Based on this information, what is the molar enthalpy of neutralizing nitric acid.

Answers:

1. a) -4.77 °C b) m = 23 grams

2. a) $0.5\ Mg + \frac{1}{16} S_8 + o_2 \rightarrow \frac{1}{2} MgSO_4 + 642.k\ kJ$

b) $1096\ kJ + 2\ BaO \rightarrow 2\ Ba + O_2$

3. a) 10 mol of b) 1.3×10^2 kJ is absorbed

4. $H_2O + D_2L_3 \rightarrow 2\ DH + L_3O + 605.25\ kJ$

5. a) 1.7 °C b) -147 kJ/mol

And MORE Questions:

1. When a sample of ethanol evaporates, it absorbs energy. If, as the ethanol evaporates, 1500 mL of water in a calorimeter goes from a temperature of 34.8°C to 3.21°C, find the energy lost by the water. (-199 kJ)

2. The combustion of a small amount of butane causes the temperature of 575 mL of water to increase from 35.9°C to 67.6°C. Calculate the heat produced by this reaction. (76.4 kJ)
3. A student is required to heat 4.20 kg of water from 2.90°C to 93.8°C. Determine the enthalpy change of the water. (1.60 MJ)

4. When 200 J of energy is removed from a beaker of water, the temperature drops from 22.0°C to 20.9°C. Find the volume of water in the beaker. (43.4 mL)

5. When 600 kJ of energy are added to 2.10 L of water with an initial temperature of 19.9°C, find the final temperature of the water. (88.1°C)

6. What is the enthalpy change when 500 mL of water in a copper kettle is heated from 22.0°C to 90.0°C? (142 KJ)

7. If 188.0 J of energy are removed from a 37.3 g sample of aluminium, find the initial temperature of the metal if the final temperature of the metal is 110°C. (116°C)

8. An unknown metal with a mass of 0.50 kg warms from 75°C to 125°C when 5.7 kJ of energy are added. Determine the specific heat capacity of the metal. Which metal is most likely used here.

   (0.23 $\frac{J}{g°C}$ Tin)

9. What quantity of heat must be supplied to heat 120 mL of water from 50.0°C to 70°C. Give your answer in MJ. (0.101 MJ)

10. A 9.40 g piece of aluminium requires 305 J of energy to raise the temperature from 85.2°C to 121.9°C. Find the specific heat capacity of the aluminium using the data provided (0.884 $\frac{J}{g°C}$)

11. Using the experimental value for the specific heat capacity of aluminium from question 10, calculate the % error. (1.45 % error ➜ shortage)