Topic 5.1   Exothermic and Endothermic Reactions

Read pages 189- 194 of your IB Chemistry text book and the resource material provided in this packet. Answer questions 1-14.

Heat and Temperature
Often the concepts of heat and temperature are thought to be the same, but they are not.

Temperature is a number that is related to the average kinetic energy of the molecules of a substance.

A regular thermometer uses the expansion of a fluid to measure temperature. When the liquid (mercury or alcohol) in a thermometer is heated the average kinetic energy of the liquid particles increases, causing the particles to take up more space expanding them up the tube.

The absolute temperature or Kelvin scale is an artificial temperature scale. The Celsius scale is based on the behavior of water molecules, with $0^{\circ}C$ being the freezing point or the point where the motion of the water molecules ceases. The Celsius scale has limited use when describing the motion of many substances, especially gases whose motions can cease at much lower temperatures. The mathematical conversion between $^{\circ}C$ and Kelvin is:

$$^{\circ}C + 273 = K$$
$$K - 273 = {}^{\circ}C$$

If Temperature is measured in Kelvin, then it is directly proportional to the average kinetic energy of the particles. In other words if you double the Kelvin temperature of a substance, you double the average kinetic energy of its molecules.

$$KE \; \alpha \; \text{Absolute temp (K)}$$

Heat is a measurement of the total energy in a substance. That total energy is the sum of the kinetic (motion) and potential (stored) energies of the molecules. It is measured in Joules (J).

The graph alongside shows how the temperature changes when a solid is heated.

Initially, the solid's temperature increases with time because the heat absorbed is used to increase the average kinetic energy of the solid particles. Since the temperature is proportional to the average kinetic energy the temperature of the solid increases.
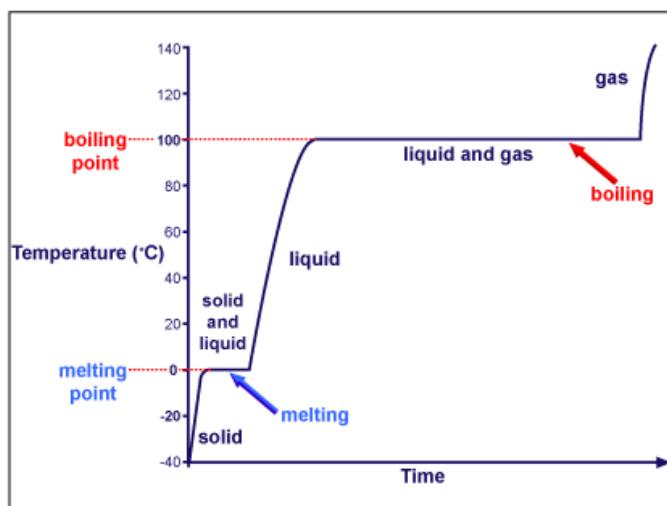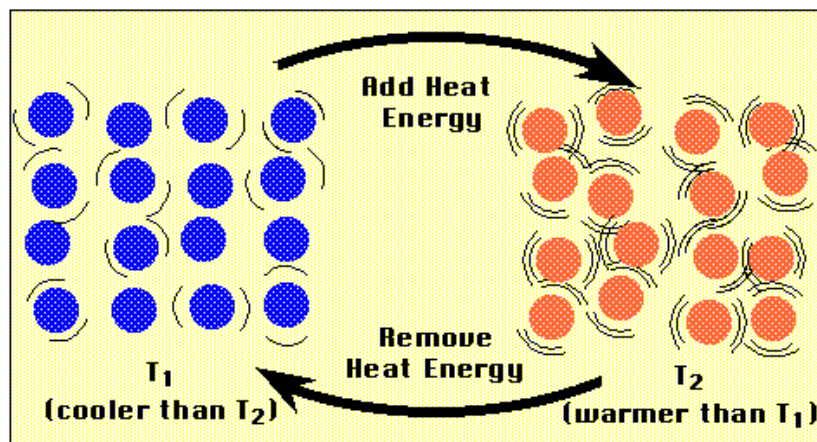
When the solid is at it melting point, it is changing state from a solid to a liquid. The heat energy is still being absorbed but is not used to increase the average kinetic energy but to increase the chemical potential energy stored in the bonds, causing them to vibrate and break. Because the average kinetic energy of the particles does not increase, the temperature does not change. The same thing happens when a liquid is at its boiling point, the temperature at which it changes state from a liquid to a gas.
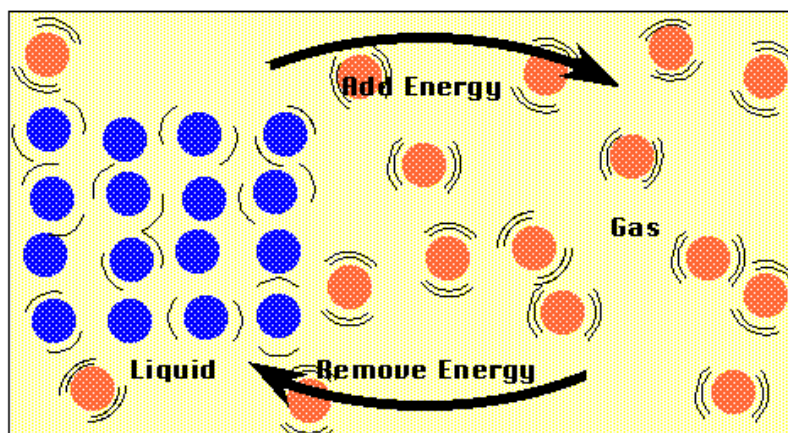
So, when heat energy is absorbed by a substance, the energy can be used to increase the average kinetic energy of the molecules, causing an increase in temperature. Alternatively the energy can be used to increase the potential energy of the molecules causing a change in state that is not accompanied by an increase in temperature. Since heat is a measure of the total energy of a substance the amount of heat depends on the moles of the substance present.
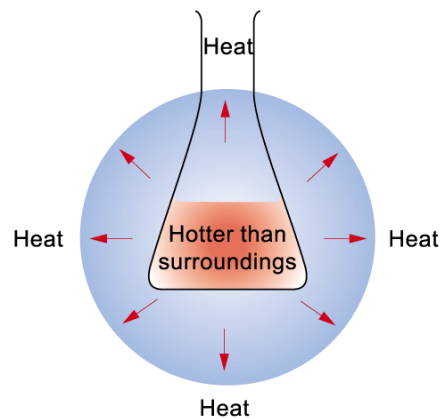
Exothermic and endothermic reactions

When a chemical reaction happens, there can be a net transfer of energy from the reactants into the surroundings (e.g. air, water is in solution) or a net transfer of energy from the surroundings to the reactants causing a change in average kinetic energy of the particles in the surrounding and a proportional change in the temperature.

Exothermic reactions

In these reactions energy is **transferred** from the reaction system **into the surroundings**. The average kinetic energy of the surroundings increases causing the reaction mixture to increase in temperature and get warmer. Some examples of exothermic reactions are:

- Burning (combustion)
- Neutralization reactions between acids and alkalis
- The reaction between water and calcium oxide

Endothermic reactions

In these reactions energy is **absorbed from the surroundings**. The average kinetic energy of the surrounding decreases causing the reaction mixture to get colder. Some examples of endothermic reactions are:

- the reaction between barium hydroxide and ammonium chloride
- the reaction between ethanoic acid and sodium carbonate
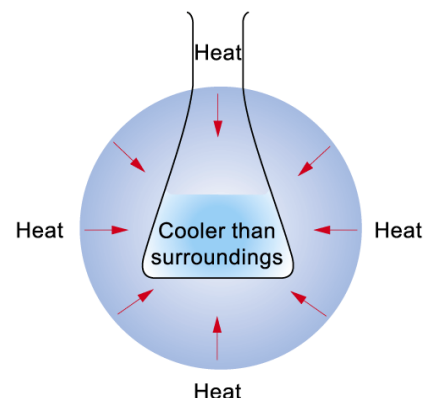
Exothermic and endothermic reactions are used extensively in everyday life and in industry. Airbags, a safety device in modern cars, utilize an exothermic reaction.

An exothermic reaction is responsible for the inflation of air bags in cars.
http://www.youtube.com/watch?v=dZfLOnXoVOQ&feature=related

Simulation showing the energy changes associated with dissolving different ionic salts in water
http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/thermochem/heat_soln.html

Cold and Hot packs used to treat muscular injury involve an endothermic reaction.

Endothermic Changes (Reactants absorb heat energy from the surroundings for the change. Average kinetic energy of particles increases)	Exothermic Changes (Reactants release heat energy into the surroundings for the change. Average kinetic energy of particles decreases)
ice melting (s → l)	combustion (burning in O_2) $CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O$
photosynthesis	respiration
Evaporation (l → g)	metal + acid → salt + hydrogen $Mg_{(s)} + 2\,HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
Melting (s → l)	Acid dissolving in water $H_2SO_{4(l)} + H_2O_{(l)} \rightarrow 2H^+_{(aq)} + SO_4^{2-}_{(aq)}$
Boiling (l → g)	Condensation (g → l)
Bond breaking	Bond making
Ammonium nitrate dissolving $NH_4NO_{3(s)} + H_2O_{(l)} \rightarrow NH_4^+_{(aq)} + NO_3^-_{aq)}$	Sodium hydroxide dissolving in water $NaOH_{(s)} + H_2O_{(l)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$
Decomposition reactions $H_2O_{(g)} \rightarrow H_{2(g)} + \tfrac{1}{2}\,O_{2(g)}$ $N_2O_{(g)} \rightarrow N_{2(g)} + \tfrac{1}{2}\,O_{2(g)}$ $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$	Neutralization reactions (acid + base → salt + water) $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(g)}$
	Forming an ionic solid from gaseous ions $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$
Sublimation (s → g) $I_{2(s)} \rightarrow I_{2(g)}$	Freezing / solidification (l → s)
Ionization $Na_{(g)} \rightarrow Na^+_{(g)} + e^-$	Calcium chloride dissolving in water $CaCl_{2(s)} + + H_2O_{(l)} \rightarrow Ca^{2+}_{(aq)} + 2Cl^-_{(aq)}$
	Magnesium sulphate dissolving in water $MgSO_{4(s)} + + H_2O_{(l)} \rightarrow Mg^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$

Increase in kinetic energy - energy absorbed - endothermic
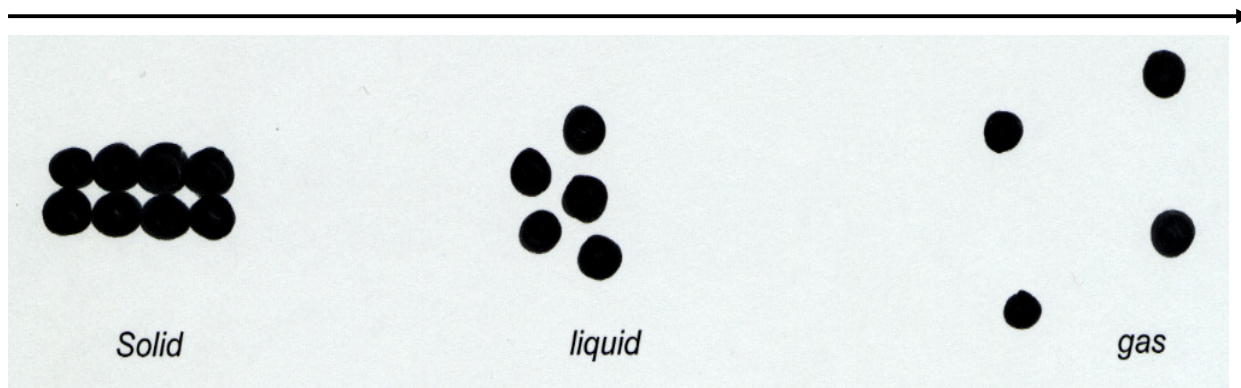
Decrease in kinetic energy - energy released – exothermic

Energy level diagrams

Chemical Potential Energy
The chemical potential energy stored in the bonds gives us a measure of a substances energy level. The higher the energy, the more chemical energy is stored in its bonds. The **reactants** and **products** in a chemical reaction usually have different energy levels, which are shown in a energy level diagram. The vertical axis on this diagram represents the energy level and the horizontal axis represents the progress of the reaction from reactants to products.

**Energy level diagrams for exothermic reactions**
In an exothermic reaction, **reactants have more energy than the products**. The difference between these two energy levels is the energy released to the surroundings, shown as a vertical drop from a higher to a lower level. Because the reactants have more energy than the products they are less stable.
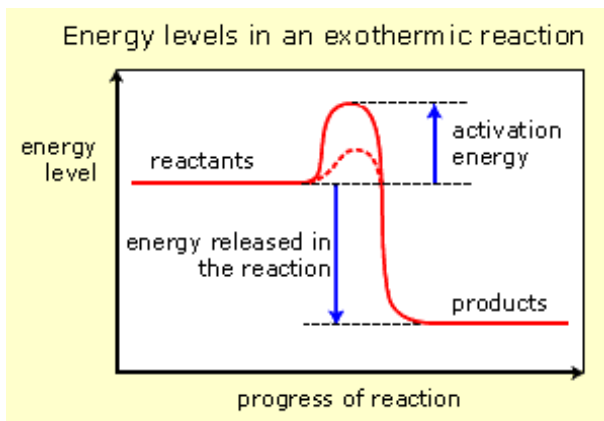
Usually some extra energy is needed to get the reaction to start. The minimum amount of energy that needs to be absorbed in order for the reactants to be converted into products is called the **activation energy**. It is drawn in energy level diagrams as a hump. **Catalysts** reduce the activation energy needed for a reaction to happen - this lower activation energy is shown by the dotted red line in the diagram here.

Energy level diagrams for endothermic reactions
In endothermic reactions the **reactants have a less energy than the products**. The difference between these two energy levels is the energy absorbed from the surroundings. It is represented in an energy level diagram as a vertical jump from a lower to a higher level - the bigger the difference, the more energy is gained. Because the reactants have less energy than the products they are more stable.

Energy levels in an endothermic reaction
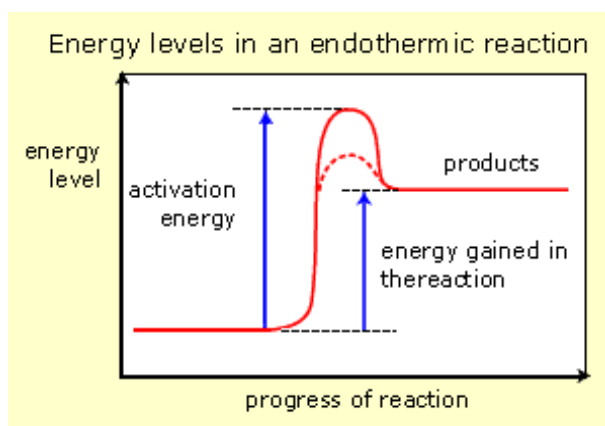
5.1 Questions

1. Define the terms exothermic and endothermic reaction.

2. Distinguish between heat and temperature.

3. Classify each of the following reactions as either exothermic or endothermic.
   a) $2H_2O(l) + heat \rightarrow 2H_2(g) + O_2(g)$
   b) $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s) + heat$

4. The complete combustion of acetic acid ($HC_2H_3O_2$) in oxygen gas to form water and carbon dioxide at constant pressure releases 871.7 kJ of heat per mole of acetic acid.
   a) Write a balanced chemical equation for this reaction.
   b) How much heat (kJ) would be released if you burned 2.0 moles of acetic acid?
   c) Draw an energy level diagram for the reaction.

5. Draw an energy level diagram for a reaction in which the total energy of the reactants is 50 kJ mol^{-1}, the total energy of products is 120 kJ mol^{-1} and the activation energy for the forward reaction is 120 kJ mol^{-1}. Label the diagram clearly. Is this reaction exothermic or endothermic?

6. When solid sodium hydroxide is dissolved in water, the temperature of the solution formed rapidly increases.
    a) Compare the total energy of the solid NaOH with that of the solution and state which is greater.
    b) Classify this reaction as endothermic or exothermic.

7. Consider the reaction A + 2B → C
   In this reaction, the total energy of the reactants is 80 kJ mol^{-1}, the total energy of the products is -90 kJmol^{-1} and the activation energy for the forward reaction is 120 kJ mol^{-1}.
    a) Draw a diagram of the energy profile for this reaction. Label the diagram.
    b) State whether the reaction is endothermic or exothermic.
    c) Calculate the energy difference between the reactants and the products.
    d) Deduce the sign of the enthalpy change.
    e) Identify with a reason, which is more stable, the reactants of products.

8. (N04/S/2)

 a) State why enthalpies of combustion reactions are negative. [1]

 b) Define activation energy / enthalpy. [1]

 c) Draw a labeled enthalpy level diagram for an exothermic and endothermic reaction showing the activation energy, E_a and enthalpy change. [4]

9. (M05/S/2) In a neutralization reaction 50 cm^3 of a 0.50 moldm^{-3} solution of sodium hydroxide is mixed rapidly in a glass beaker with 50 cm^3 of a 0.050 moldm^{-3} solution of sulfuric acid.
                    Initial temperature of each solution = 19.6°C
                    Final temperature of the mixture = 23.1°C
 a) Write an equation for the reaction. [1]
 b) State with a reason whether the reaction was exothermic or endothermic. [1]

10. (M04/S/2) State the conditions under which standard enthalpy changes are measured. [1]

11. Consider two beakers of water. Both have the same temperature, but the 100 cm^3 of water contains twice as much heat as the 50 cm^3. Explain why?

12. (M00/S/2)
 a) Draw an enthalpy level diagram for a neutralization reaction.
      i)   Indicate on your diagram the enthalpy change of the reaction and deduce its sign.
      ii)   Compare the relative stabilities and strengths of the bonds of the reactants and products. [4]
      iii)   Define the term standard enthalpy change of a reaction. [1]

13. Develop an argument that involves diagrams to explain which contains more energy – a swimming pool of cold water or a pot of boiling water.

14. Complete the sentences in the table below [5]

Exothermic	Endothermic
ΔH sign is _____	ΔH sign is _____
Energy _____	Energy _____
Products more stable than _____ (as energy decreases stability increases)	Reactants more stable than _____
Products have less energy than _____	Products have more energy than _____
Temperature _____	Temperature _____