

• TOPIC-1 THE HIGHS AND LOWS OF ENERGY

In this chapter, we will study the energy associated with chemical reactions. This energy is usually in the form of heat flowing out of or into a physical or chemical reaction. We'll perform a series of experiments and carefully note the type of phenomena (physical or chemical), and the direction of heat transfer (absorbed or released)

We can classify three types of changes. For each change, there is an *exchange* of energy either to or from the environment. The changes are

- PHYSICAL CHANGE (involves the least amount of energy)
- CHEMICAL CHANGE (involves an intermediate amount of energy)
- NUCLEAR CHANGE (involves the most amount of energy)

Recall that in a **physical change** a substance only changes its physical appearance or physical state. The substance is the same before and after the change. In effect, the molecules of a substance in a physical change always remain intact and do not break up. Hence, relatively little energy is involved when physical changes occur.

In a **chemical change**, the molecules break into atoms. This breaking of molecular bonds requires energy. Once the energy is supplied, the atoms rearrange themselves forming new bonds and producing new substances. During the process of forming new bonds, energy is released. Because bonds are broken and new ones formed, *there is much more energy associated with chemical changes than with physical changes*. While the breaking of bonds requires energy, the forming of new bonds releases energy. The net difference between the energy required for bond breaking and the energy released in bond forming determines whether the overall reaction gives off or absorbs energy.

Chemical reactions which **give off heat** are called **exothermic** reactions. Chemical reactions which **absorb heat** are called **endothermic** reactions.

To remember that an exothermic reaction gives off heat and an endothermic reaction absorbs heat, note that the root "endo" means "inside" while the root "exo" means "outside" such as in the word "exit".

In a nuclear change, there is a rearrangement of the particles making up the atom. As a result, there is a great deal of energy involved when a nuclear reaction occurs.

• EXPERIMENT-1 Hot and cold reactions

Purpose: To investigate heat transfers.

This experiment consists of observing six phenomena. During each phenomenon, you will make quantitative observations. You will measure the initial and final temperatures of the solutions. Then, you will determine whether there has been a *transfer* of heat energy to or from various physical and chemical phenomena.

Phenomenon 1: Moisten a Q-tip with alcohol and spread it on the back of your hand.

For the following phenomena, measure the initial temperature of the water and then measure either the highest or the lowest temperature of the solution.

Phenomenon 2: Add 3 mL to a test tube with water then add some ice to the water.

Phenomenon 3: Add 3 mL to a test tube with water then, using the tip of a spatula, add a tiny amount of "Drano" ($\text{NaOH}_{(s)}$) to the water.

Phenomenon 4: Fill a test tube with 3 mL water then add 3 mL of a 1.0 M sodium hydroxide ($\text{NaOH}_{(aq)}$) solution.

Phenomenon 5: Fill a test tube with 3 mL water then add a drop of phenolphthalein. Next, add 3 mL of 1.0 M sodium hydroxide solution ($\text{NaOH}_{(aq)}$). Finally, add 3 mL of 1.0 M hydrochloric acid solution ($\text{HCl}_{(aq)}$).

Phenomenon 6: Fill a test tube with 3 mL water then, using the tip of a spatula, add a tiny amount of sodium nitrate ($\text{NaNO}_{3(s)}$) to the water.

Write your observations in the table below.

Phenomenon	Qualitative Observation	Initial T (± °C)	Final T (± °C)
1			
2			
3			
4			
5			
6			

Analyze the results of your experiments by filling-in the following table with a check mark (✓) in the appropriate columns.

Phenomenon	Reactants	Change		Heat Behaviour		
		Physical	Chemical	Colder	Warmer	Stable
1	Alcohol _(l)					
2	Water _(l) + Ice _(s)					
3	Water _(l) + Drano _(s)					
4	Water _(l) + NaOH _(s)					
5	NaOH _(aq) + HCl _(aq)					
6	Water _(l) + NaNO _{3(s)}					

Classify the experimental phenomena as physical or chemical changes according to the flow of heat energy.

	Endothermic Phenomenon	Exothermic Phenomenon	Other Phenomenon
Physical Change			
Chemical Change			

1) Alcohol creates a cooling sensation on the skin. Explain why.

◆End of experiment

2) We will now analyze the steps involved in the combustion of a candle from the point of view of energy flow

Exothermic Endothermic

- | | | | |
|---|--------------------------|--------------------------|--|
| a) A flame is brought close to the wick of a candle absorbing energy from the flame. | <input type="checkbox"/> | <input type="checkbox"/> | |
| b) The wick of the candle burns with a flame giving out heat energy | <input type="checkbox"/> | <input type="checkbox"/> | |
| c) The heat from the flame melts the wax of the candle. | <input type="checkbox"/> | <input type="checkbox"/> | |
| d) By capillary action, the liquid wax comes into contact with the hot flame and vaporizes. | <input type="checkbox"/> | <input type="checkbox"/> | |
| e) The vaporized wax burns with a flame thereby producing heat and light. | <input type="checkbox"/> | <input type="checkbox"/> | |

3) When a candle burns, it emits heat and light. Where does this energy come from? Explain.

• TOPIC-2: ENERGETICALLY YOURS

When we want to cool a soft drink, we add a few ice cubes. If a cup of coffee is too hot, we may add milk. When bath water is too cold, we turn on the hot water faucet to add hot water.

From experience, we know that the temperature of one substance affects the temperature of another substance. What exactly happens when two substances at different temperatures come into contact? What molecular activities go on? Why is the influence mutual?

It is important not to confuse heat with temperature. Whereas **heat is energy, temperature is a comparison between two different levels of heat energy**. If we know the temperature of a material, we only know if it is hotter or colder in comparison to another material. We do not know, however, the amount of heat the material has (that is, its heat content).

To obtain a better understanding of the difference between temperature and heat, we will use an analogy comparing heat energy with money. By way of analogy, the more money we have, the more heat energy we have and vice-versa.

Consider the following analogy

As you know, a substance consists of molecules. If we consider a group of people as a substance, then each person in the group is like a molecule in the substance; the greater the amount of substance, the greater the number of molecules. Likewise, the larger the group, the greater would be the number of people making up the group.

Assume a group of 100 people. How much money (pocket money) does each person carry? Suppose we ask each person to tell us the amount of money they carry. Adding each individual amount, we obtain a total of, say, \$4000. We can now calculate the average amount of money for each person. In this case, the average turns out to be \$40 ($4\ 000/100$). An average of \$40 is the same as if every person in the group has exactly \$40. Of course, because it is an average, some people will have more than \$40 and some people will have less than \$40. Analogously, this is equivalent to a temperature of $400\ ^\circ\text{C}$ and a total heat content of 4 000 heat units. In reality, the total heat content of a material depends upon three things

- The nature of the substance (type of substance)
- The temperature of the substance (average energy of the molecules)
- The amount of substance (mass)

Remember, knowing only the temperature of a substance, we cannot determine its heat content.

As an example, consider two substances; substance-A and substance-B. By way of analogy, these two substances can be considered as two groups of people (group-A and group-B). Suppose group-A consists of, say, 10 people each with an average amount of \$1 000, and group-B consists of, say, 20 000 people each with an average of \$ 1. Hence, the total amount of money for group-A is \$ 10 000 and the total amount for group-B is \$20 000. In comparing the two groups, it is obvious that the average person in group-A is richer (\$1 000 average) than the average person in group-B (\$1 average). Collectively, however, group-B has more money (\$20 000) than group-A (\$ 10 000).

Analogously, suppose we have a bathtub full of water at, say, 200. °C and a glass full of water at 800. °C. The average energy of the water molecules in the bathtub is, therefore, 20 units of heat energy as compared with an average of 80 units of heat energy for each water molecule in the glass. However, the heat content of the water in the bathtub is more than the heat content of the water in the glass. This is simply because there are more molecules of water in the bathtub than in the glass.

Temperature, therefore, tells us the average heat energy of individual molecules whereas heat tells us the total heat energy of a substance. If the glass of water at 80° comes into contact with the bath-tub water at 20°, heat energy will flow from the glass to the bath-tub even though the water in the bath-tub already contains more heat energy than the water in the glass.

Note : The direction of heat flow is easily determined by temperature. When two substances at different temperatures come into contact with each other, heat always flows from the substance with the higher temperature to the substance with the lower temperature, regardless of the quantity of either substance. Temperature, however, does not tell us the quantity of heat flow.

The first law of energy states that "**Energy cannot be created nor destroyed**". That is, in any reaction (physical or chemical), energy is conserved. This means that a system always keeps the same amount of energy it has. It cannot gain or lose energy by itself. To lose energy, a system must release (give out) energy to its surroundings. Conversely, to gain energy, a system must absorb (take in) energy from its surroundings.

Before we state a mathematical equation concerning heat flow, you should note that when a system gains energy, the action is "positive" as the system increases its energy content. Conversely, when a system loses energy, the action is "negative" as the system decreases its energy content.

Remember: An energy **loss** is *negative* as energy must be subtracted from a system's total energy content. On the other hand, a **gain** in energy is *positive* as energy is added to a systems total energy content.

When two substances (systems) at different temperatures come into contact with each other, heat energy will flow from the material at the higher temperature to the material at the lower temperature until both materials reach the same temperature. This means the hotter material will lose heat energy and the colder material will gain heat energy. The hotter material is said to have a "negative heat flow" (energy loss) while the colder material is said to have a "positive heat flow" (energy gain).

Thus, we may write the following "word equation"

$$\text{heat lost} = \text{heat gained}$$

In the word equation above, note that the words "heat lost" implies a negative sign. That is, the negative sign is understood by the nature of the statement. Likewise, the positive sign is understood by the nature of the statement "heat gained". We need not include the signs "positive" or "negative" when stating the heat flow equation in words.

In writing the word equation mathematically, however, signs must be used to designate the implied "negative" or "positive" meaning. It is customary to use capital letter Q to represent the quantity of heat. And, to differentiate a loss from a gain in the quantity of heat, we use the minus and plus signs respectively. Thus, the word equation becomes

$$\text{heat lost} = \text{heat gained}$$

or

$$-Q_{\text{lost}} = +Q_{\text{gained}}$$

The formula used to calculate the quantity of **heat flow** to or from a substance is

$$Q = (\text{mass}) \times (\text{specific heat capacity of substance}) \times (\text{temperature change})$$

where

- mass is the mass of the substance (in grams)
- specific heat capacity is the amount of heat ($\text{J/g}\cdot^{\circ}\text{C}$) the substance can absorb or release
- change in temperature is the difference between the final minus the initial temperature

Using the delta notation to represent the words "a change in temperature", we may write the "change in temperature" as ΔT .

Note that $\Delta T = T_f - T_i$ (final temperature minus initial temperature)

Thus, using m for mass, c for specific heat capacity, and ΔT for the change in temperature, the heat formula can be written as follows

$$Q = mc\Delta T$$

or

$$Q = mc(T_f - T_i)$$

• Heat Energy ($Q=mc\Delta T$) Problems

1. Many hot water heaters use the combustion of natural gas to heat the water in the tank. When 150.0 L of water at 10.0 °C is heated to 65.0 °C, how much energy flows into the water?
2. In an industrial plant, 100.0 kg of steam is heated from 100.0 °C to 210.0 °C. The specific heat capacity of steam is 2.01 J/g·°C. Calculate the quantity of heat that flows into the steam.
3. Some North American native peoples use the rocks heated in fire pits to produce steam in ‘sweat lodges’ for purification rites. If the specific heat capacity of the rocks is 0.86 J/g·°C, what quantity of heat is released by a 2.5 kg rock as it cools down from 350.0 °C to 15.0 °C?
4. Exactly 500.0 kJ of heat is absorbed at a constant pressure by a sample of helium. The temperature increases by 15.0 K. The mass of the helium sample is 6.42 kg. Find the specific heat capacity of the helium.
5. How many kilojoules of heat are required to raise the temperature of 165 mol of water from 10.5 °C to 47.5 °C?
6. Large beds of rocks are used in some solar-heated homes to store heat. Calculate the quantity of heat absorbed by 50.0 kg of rocks if their temperature increases by 12.0 °C. The specific heat of the rocks is 0.86 J/g·°C.
7. If the rocks in question 6 absorbed 500. kJ of heat and their initial temperature was 18.0 °C, what would their final temperature be?

The Specific Heat of a Metal

The amount of heat energy that is required to raise the temperature of one gram of a substance by one degree Celsius is called the specific heat capacity, or simply the specific heat, of that substance.

Objective: To determine the specific heat of lead ($\text{Pb}_{(s)}$).

Procedure:

1. Record the mass of lead shot, $\text{Pb}_{(s)}$ in data table.
2. Transfer the lead shot to a large, dry test tube. Be careful to pour the lead shot into the tube slowly so that the bottom of the test tube is not broken in this process. Suspend the test tube in the boiling water with a utility clamp. Position the test tube so that the lead shot is below the level of water in the beaker. (See Figure 1.) Allow the test tube to remain in the boiling water bath for at least 10 minutes. Proceed to Step 4 while the lead shot is heating.

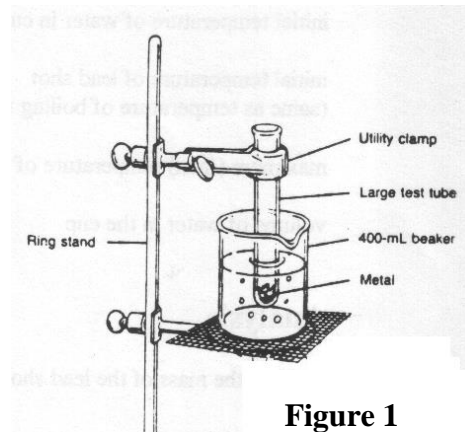


Figure 1

3. Carefully measure out 100.0 mL of water in a graduated cylinder, and pour the water into a styrofoam cup. Place the cup in a 250-mL beaker for support. (See Figure 2)
4. After the lead shot has been heating for at least **10 minutes**, record the temperature of the water in the styrofoam cup and the $\text{Pb}_{(s)}$ shot in the boiling water bath.

5. Remove the test tube from the bath, using the clamp as a holder. Carefully, but quickly, pour the $\text{Pb}_{(s)}$ shot into the water in the styrofoam cup. Cover the Styrofoam cup with the plexiglass lid. Place a thermometer through the hole in the plexiglass lid. Use the thermometer to gently stir the lead shot. Observe the temperature frequently. As the temperature begins to change more slowly, watch the thermometer continuously so as not to miss the maximum (final) temperature reached. Record this maximum (final) temperature in data table.

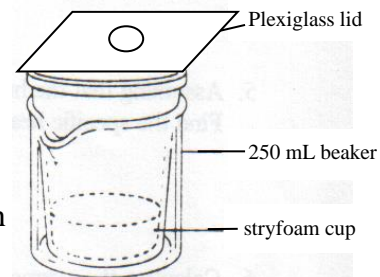


Figure 2: Calorimeter

6. Decant the water off the lead shot and return the metal to paper towel in weighing dish to be dried.

Data Table

Title: _____

Mass of lead shot	g
Volume of water in the cup	mL
Initial temperature of water in calorimeter	°C
Initial temperature of lead shot (same as temperature of boiling water)	°C
Maximum(final) temperature of water + lead shot	°C

Analysis: (*Show calculations with appropriate units.*)

1. Determine the mass of the water in the cup. Assume the density of water is 1.000 g/mL.
2. Calculate the change in temperature of the water (ΔT_{water}) and of the lead shot (ΔT_{lead}).
3. Find the amount of heat energy gained by the water.
4. Assuming that the heat energy gained by the water is equal to the heat energy lost by the lead, find the specific heat capacity of the lead.

Conclusion:

Using the accepted value of 0.159 J/g·°C, calculate the % error.

What is (are) the source(s) of this error?

• Specific Heat Problems

1. Which requires more energy warming 15.0 g of water or warming 60.0 g of aluminum from 25.0 °C to 37.0 °C? (specific heat values: Al = 0.899 J/g·°C, H₂O = 4.19 J/g·°C)
2. How much heat energy in kilojoules is required to heat all the aluminum in a roll of aluminum foil (500.0 g) from room temperature (25.0 °C) to the temperature of a hot oven (250.0 °C)?
3. Ethylene glycol (CH₂OH)₂ is often used as an antifreeze in cars. It has a specific heat capacity of 2.42 J/g·°C.
 - a) Which requires more heat energy to warm from 25.0 to 100.0 °C, pure water or an equal mass of pure ethylene glycol?
 - b) If the cooling system in an automobile has a capacity of 4.73 L, compare water and ethylene glycol as to the quantity of heat energy the liquid in the system absorbs on raising the temperature from 25.0 to 100.0 °C. Assume the density of ethylene glycol is 1.11 g/mL.
4. One way to cool a cup of coffee is to plunge an ice-cold piece of aluminum into it. Suppose a 20.0 g piece of aluminum is stored in the refrigerator at 0.0 °C and then dropped into a cup of coffee. The temperature of the coffee drops from 90.0°C to 75.0 °C. How much heat energy in kilojoules did the aluminum absorb?
5. A piece of iron (400.0 g) is heated in a flame and then plunged into a beaker containing 1.00 kg of water. The original temperature of the water was 20.0 °C but it is 32.8 °C after the iron is dropped in. What was the original temperature of the hot iron? ($c_{\text{Fe}} = 0.444 \text{ J/g}\cdot^{\circ}\text{C}$)
6. A 192 g piece of Cu_(s) was heated in a boiling water bath and then dropped into a beaker containing 750.0 mL of water at 4.0 °C. What is the final temperature of the Cu_(s) and water after they come to thermal equilibrium? ($c_{\text{Cu}} = 0.385 \text{ J/g}\cdot^{\circ}\text{C}$)

- **IMPORTANT:** It is important to realize that the heat formula **cannot** be used to find the heat content of a substance. The heat formula tells us only the quantity of heat flowing into or flowing out of a substance. That is, the formula tells us the **change**, increase or decrease, in the heat content of a substance.

We can simplify the use of the heat equation when mixing two like liquids at different temperatures such as water with water. We will adopt the following convention:

T = final temperature of mixture
 T_h = temperature of hot water
 T_c = temperature of cold water
 m_h = mass of hot water
 m_c = mass of cold water

since:

heat lost = heat gained

$$-Q_h = +Q_c$$

and

$$-m_h c \Delta T_h = +m_c c \Delta T_c \quad (\text{Cancel } Ac \cong \text{ on both sides when liquids are alike})$$

$$-m_h(T_f - T_i) = +m_c(T_f - T_i)$$

or

$$-m_h(T - T_h) = m_c(T - T_c) \quad (\text{Substitute symbols as defined above})$$

$$-m_h T + m_h T_h = m_c T - m_c T_c$$

$$-m_h T - m_c T = -m_h T_h - m_c T_c \quad (\text{Bring terms with `T` on left side})$$

$$m_h T + m_c T = m_h T_h + m_c T_c \quad (\text{Divide all terms by } -1)$$

$$T(m_h + m_c) = m_h T_h + m_c T_c$$

Thus:

$$T = \frac{m_h T_h + m_c T_c}{m_h + m_c}$$

Example-1

Suppose we mix 200.0 mL of hot water at 85.0 °C with 30.0 mL of cold water at 5.0 °C. What is the temperature of the mixture? (Note: 1 mL of water = 1 g)

Using the above formula, we have:

$$T = \frac{m_h T_h + m_c T_c}{m_h + m_c}$$

$$T = \frac{[(200.0 \text{ g})(85.0 \text{ }^\circ\text{C}) + (30.0 \text{ g})(5.0 \text{ }^\circ\text{C})]}{(200.0 \text{ g} + 30.0 \text{ g})}$$

$$\mathbf{T = 75 \text{ }^\circ\text{C}}$$

Example-2

Suppose we mix 500.0 mL of hot water at 60.0 °C with 50.0 mL of cold water at 0.0 °C. What is the temperature of the mixture?

$$T = \frac{m_h T_h + m_c T_c}{m_h + m_c}$$

$$T = \frac{(500.0)(60.0) + (50.0)(0.0)}{500.0 + 50.0}$$

$$\mathbf{T = 55 \text{ }^\circ\text{C}}$$

- 1) Find the final temperature of a mixture made by mixing 125.0 mL of water at 25.0 °C with 60.0 mL of water at 5.0 °C.
- 2) 500.0 mL of water at 45.0 °C is mixed with 600.0 g of water at 90.0 °C. Calculate the temperature of the mixture?
- 3) A beaker contains 400.0 mL of water at 60.0 °C. A student adds 250.0 mL of water which is at a temperature of 20.0 °C and then adds another 300.0 mL of water at 70.0 °C. Find the temperature of the final mixture.

• **EXPERIMENT-2: Hot and cold equals what?**

Purpose: To investigate heat flow.

- Step-1: Pour 50.0 mL of cold water into a calorimeter, close the cover, and let it stand (about 10 seconds) to allow the temperature to stabilize (between 4.0 °C to 10.0 °C).
- Step-2: Pour 50.0 mL of hot water into a calorimeter, close the cover, and let it stand (about 10 seconds) to allow the temperature to stabilize.
- Step-3: If we mix together the water in the two calorimeters, predict the temperature of the mixture. Record your prediction in the table below.
- Step-4: Pour the cold water into the calorimeter containing the hot water. Using a thermometer, read and record the temperature of the mixture.
- Step-5: Pour 75.0 mL of cold water into an empty calorimeter, close the cover, and let it stand (about 10 seconds) to allow the temperature to stabilize.
- Step-6: Pour 25.0 mL of hot water into an empty calorimeter, close the cover, and let it stand (about 10 seconds) to allow the temperature to stabilize.
- Step-7: Predict the temperature of the mixture if the contents of both calorimeters were mixed together and record your prediction in the table below.
- Step-8: Pour the cold water into the calorimeter containing the hot water and measure the temperature of the mixture with a thermometer. (Record the value in the table below.)
- Step-9: Pour 75.0 mL of hot water into an empty calorimeter, close the cover, and let it stand (about 10 seconds) to allow the temperature to stabilize.
- Step-10: Pour 25.0 mL of cold water into an empty calorimeter, close the cover, and let it stand about 10 seconds to allow the temperature to stabilize.
- Step-11: Predict the temperature of the mixture of both calorimeters if their contents were mixed together. (Record your prediction in the table below)
- Step-12: Pour the cold water into the calorimeter containing the hot water and measure the temperature of the mixture with a thermometer. (Record the value in the table below.)

Write the predictions and measured values below:

Prediction	Hot Water		Cold Water		Mixture		
	V (± mL)	T (± °C)	V (± mL)	T (± °C)	V (mL)	Predicted T (°C)	Measured T (°C)
1	50.0		50.0				
2	25.0		75.0				
3	75.0		25.0				

◆ Remember: When two substances at different temperatures come into contact, heat will flow from the hotter to the colder substance. The substance that loses heat has a **negative Q** value while the substance that gains heat has a **positive Q** value. Heat given off is **negative** while heat gained is **positive**.

1) What are the two factors that influence the temperature of the mixture?

End of experiment

• Topic 3 - Hidden Energy

Matter possesses internal energy. For example, at the moment of combustion, gasoline releases an enormous amount of "hidden" energy. The energy is "hidden" in that it cannot be seen simply by looking at the liquid gasoline.

Recall the burning candle? It released two kinds of "hidden" energy, heat and light. Food also contains "hidden" energy. During the process of digestion, a kind of slow combustion or burning, energy is released and absorbed by the body.

Often during a chemical reaction such as combustion, digestion, etc., energy is released. As you know, such reactions are called *exothermic* reactions.

Some chemical reactions absorb energy. For example, in photosynthesis, matter absorbs light energy and stores it in its molecules. The stored energy is "hidden" energy. Some of this hidden energy is liberated when, for example, we eat spinach. In digesting the spinach, the hidden (or stored) energy is released to the body. Such reactions are called *exothermic* reactions.

• EXPERIMENT-3: Absorbing and releasing energy

Purpose: To investigate hidden energy.

In this experiment we will perform some chemical reactions and determine the quantity of heat energy absorbed or released during the process. Specifically, we will investigate the dissolving of solutes in water.

Prelab: Calculate the molar mass of:

Group A		Group B	
Formula	Molar Mass (g/mol)	Formula	Molar Mass (g/mol)
$\text{NaC}_2\text{H}_3\text{O}_2(\text{s})$		$\text{NH}_4\text{NO}_3(\text{s})$	
$\text{NaOH}(\text{s})$		$\text{KNO}_3(\text{s})$	
$\text{KOH}(\text{s})$		$\text{NaNO}_3(\text{s})$	

GROUP-A Procedure: Dissolving a solid in water

- Step- 1: Select a chemical from Group-A in the Prelab section above and write its name and chemical formula in the table below (solute).
- . Step-2: Calculate and record the molar mass of your chosen chemical in the table below
- Step-3: Mass 0.05 mole of your chemical (mass = moles x molar mass) into calorimeter.
- Step-4: Using a thermometer, read and record the temperature of the tap water.
- Step-5: Measure 50.0 mL of tap water and pour it into a calorimeter (solvent).
- Step-6: Cover the calorimeter with the lid, which supports the thermometer.
- Step-7: Using the thermometer, gently stir the solution then record the highest or lowest temperature.

Repeat steps 1 to 7 above with another substance from Group-A and then two substances from Group-B for a total of four substances.

DATA TABLE:

Write your experimental data in the table below for **Group A solids(solutes)**.

Name of solute	mm (g/mol)	n _{solute} (mol)	m _{solute} (± g)	T _{initial} (± °C)	T _{final} (± °C)

Calculate and record the " ΔT " of the mixtures above and record any observations below:

Write your experimental data in the table below for **Group B solids(solutes)**.

Name of solute	mm (g/mol)	n _{solute} (mol)	m _{solute} (± g)	T _{initial} (± °C)	T _{final} (± °C)

Calculate and record the " ΔT " of the mixtures above and record any observations below:

1) Were the reactions you observed endothermic or exothermic? Explain:

2) For exothermic reactions, have you observed an increase in the water temperature?
If so, where did the heat energy come from?

3) For endothermic reactions, have you observed a decrease in the water temperature?
If so, where did the heat energy go?

4) During an _____ reaction, heat energy flows from the reactants to the surrounding environment.

During an _____ reaction, heat energy flows from the surrounding environment to the reactants.

◆ **End of experiment**

The change in temperature of the surrounding environment is evidence that a reaction is either endothermic or exothermic. If the temperature of the surrounding environment goes down (heat taken in), the reaction is **endothermic**. If the temperature of the surrounding environment goes up (heat given out), the reaction is **exothermic**. It is convenient to use water as the surrounding medium. Thus, by measuring the change in water temperature, we can establish the type of reaction as exothermic or endothermic.

When a fixed mass of substance dissolves in water, we observe a large change in temperature when the quantity of water is small. Yet, when the same amount of substance is dissolved in a large quantity of water, we observe a much lower change in the temperature of the water. This is normal since in both cases the same amount of heat is absorbed in the first case by a small quantity of water and in the second case by a large quantity of water.

Recall the analogy between heat and money? The small quantity of water is analogous to a small group of people. The large quantity of water is analogous to a large group of people. The heat given out is analogous to a fixed sum of money, say \$ 1 000. Thus, if we were to hand out to each group this sum of money, it is obvious that the smaller the group (less water), the greater the amount of money each individual receives (the higher the temperature). Analogously, this is equivalent to increasing the average heat energy of each molecule. The average heat energy becomes greater the smaller the quantity of water (and vice-versa). Thus, although both quantities of water receive the same amount of heat, the smaller quantity is observed to be at a higher temperature than the larger quantity.

Note that the nature of the surrounding medium is also responsible for the change in temperature. Not all mediums absorb equally. Some mediums absorb more energy than others. This means that the same amount of heat will not raise the temperature of each medium equally. This characteristic of matter is called the "heat capacity" of a substance. The greater the heat capacity, the more heat energy the substance absorbs before its temperature rises. In effect, the heat capacity tells us the amount of heat a substance must receive in order to raise its temperature one degree. Of course, in order to compare the heat capacity of different substances, we must use the same quantity of matter (like one gram).

For example, suppose we have exactly one gram of two different materials, substance-A and substance-B. Suppose further that the heat capacities of these two substances are as follows:

substance-A:	$c = 10$ heat units
substance-B:	$c = 50$ heat units

A heat capacity of 10 heat units for substance-A means we must supply to this material 10 units of energy per gram for its temperature to rise 1°C . On the other hand, to raise the temperature of substance-B by 1°C , we'll need 50 units of heat energy per gram.

Now, suppose we have 35 units of heat energy and one gram of each substance. If we give this energy to substance-A we notice a rise in temperature of 3°C and some energy left over. If we give this same energy to substance-B, we notice no rise in temperature and no energy left over. The simple reason for this is that the specific heat capacity of substance-B, designated by the letter "c", is much higher than the specific heat capacity of substance-A.

The following table lists the specific heat capacities of some substances at room temperature. The specific heat capacity of a substance will vary slightly at different temperatures. Because the variation is small, however, to keep things simple, we will assume that the heat capacity of a material does not change with temperature.

Substance	Mass (g)	Temperature change ($^{\circ}\text{C}$)	Specific heat capacity - c ($\text{J/g}\cdot^{\circ}\text{C}$)
Water	1.00	1.00	4.19
Methanol	1.00	1.00	2.55
Ethanol	1.00	1.00	2.45
Acetone	1.00	1.00	2.16

Specific heat capacity is the amount of energy required to raise the temperature of one gram of a substance by one degree Celsius.

5) Suppose we supply, in turn, exactly 100 J of heat energy to 25.0 g of water, methanol, ethanol, and acetone; which of these substances would have the greatest rise in temperature? Explain.

Let us use the heat formula to calculate the quantities of heat flow involved for the reactions we performed in experiment 3.

$$Q = mc\Delta T = mc(T_f - T_i)$$

6) Using the heat formula, calculate the quantity of heat energy for the two *exothermic* and the two *endothermic* reactions observed in the lab:

Exothermic reactions

Chemical formula of solute	n_{solute} (mol)	m_{solute} (g)	V_{solvent} (mL)	$\Delta T(^{\circ}\text{C})$	$C_{\text{H}_2\text{O}(l)}$ (J/g $\cdot^{\circ}\text{C}$)
	0.05		50.0		4.19
	0.05		50.0		4.19

Calculate the quantity of heat released by the solute to the water in kJ/g:

Endothermic reactions

Chemical formula of solute	n_{solute} (mol)	m_{solute} (g)	V_{solvent} (mL)	$\Delta T(^{\circ}\text{C})$	$C_{\text{H}_2\text{O}(l)}$ (J/g $\cdot^{\circ}\text{C}$)
	0.05		50.0		4.19
	0.05		50.0		4.19

Calculate the quantity of heat absorbed by the solute from the water in kJ/g:

◆ **Reminder:** Recall that the amount of a substance can be specified either by mass (grams) or by moles (molecules). By way of analogy, this is similar to selling food where it may be sold either by mass (kilograms) or by pieces such as per dozen. The mole specifies a specific number of molecules called an Avogadro number (6.02×10^{23}). Thus, 18.02 g of water represents one mole (mass = moles \times molar mass). Remember, both grams and moles define the same thing, a quantity of matter. Grams define a quantity of mass while moles define a quantity of mass by number (multiples of 6.02×10^{23}).

Chemists find the mole a more convenient method of specifying the mass of a substance than the gram. Consequently, when dealing with heats of reaction, rather than specify the quantities of heat **per gram**, they specify them instead **per mole**. Thus, the heat associated per mole is defined as the **molar heat**, that is the heat involved **per mole** of substance. Therefore, in a chemical reaction it is customary to express the heat absorbed or the heat released as the **molar heat of reaction** (understood "**per mole of product produced**").

The unit for the molar heat of reaction is the joule (J or kJ). The symbol used to express the change in heat is ΔH . Since an exothermic reaction is one in which the reactants release energy to the surroundings (thus diminishing their energy content), an exothermic reaction is designated as $-\Delta H$. Conversely, since an endothermic reaction is one in which the reactants take in energy from the surroundings (thus increasing their energy content), an endothermic reaction is designated as $+\Delta H$.

➤ **Remember:** An *exothermic* reaction has a **negative** ΔH value ($\Delta H < 0$).
An *endothermic* reaction has a **positive** ΔH value ($\Delta H > 0$).

Since ΔH represents a change in heat content, it is important to always include the sign (plus or minus) when writing ΔH . Indeed, a **ΔH value written without its sign is meaningless**, as we have no way of knowing whether heat was absorbed or released.

$\Delta H = H_p - H_R$ (and when $H_p > H_R$ ΔH is positive, when $H_p < H_R$ ΔH is negative)

7) Convert the heat of reaction for dissolving a solute from kJ/g to kJ/mol.

8) Enter the class average values of the molar heats of solution for all the substances studied in the table below (be sure to include the correct sign):

Name of solute	Molecular formula	ΔH (kJ/mol)

NOTE:

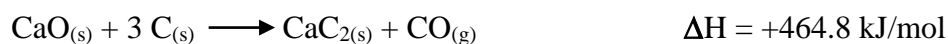
To determine the heat involved when a substance dissolves in water, we measure the heat gained or lost by the water. In doing so, however, we assume that all the heat energy goes to the water. Strictly speaking, this is not true. The solute will absorb some of the heat energy. However, since this energy is relatively small, for all practical purposes we may ignore any heat absorbed by the solute.

• Thermochemical Equations Problems

1. Energy is stored in the human body in the form of adenosine triphosphate (ATP). It is formed by the reaction between adenosine diphosphate (ADP) and phosphoric acid (H_3PO_4).



- a) Is the reaction endothermic or exothermic?
b) Rewrite the equation using ΔH notation.
2. Calcium carbide, $\text{CaC}_{2(\text{s})}$, is manufactured by reacting lime ($\text{CaO}_{(\text{s})}$) with carbon at high temperature. The carbide is then used to make acetylene.



- a) Is the reaction endothermic or exothermic?
b) Rewrite the equation with the heat energy in the equation.
4. $\frac{1}{2} \text{ H}_{2(\text{g})} + \frac{1}{2} \text{ Br}_{2(\text{g})} \longrightarrow \text{HBr}_{(\text{g})} \quad \Delta\text{H} = -36 \text{ kJ/mol HBr}$
- a) Is the reaction endothermic or exothermic?
b) Rewrite the equation using 1 mol of H_2 and put the heat energy into the equation.

5. When slaked lime is added to water in the production of cement, the reaction $\text{Ca}(\text{OH})_{2(\text{s})} \longrightarrow \text{Ca}^{2+}_{(\text{aq})} + 2 \text{ OH}^{-}_{(\text{aq})}$ takes place causing the temperature to increase.

- a) Is the reaction exothermic or endothermic?
b) What is the sign of the ΔH ?
6. A commercial hand warmer called Re-Heat consists of a plastic pouch containing a chemical dissolved in water. A small piece of flexible metal is also in the pouch. When the metal piece is flexed, crystals start to form near it and soon spread throughout the solution. At the same time, the pouch and its contents become pleasantly warm. Determine the sign of ΔH .
7. Write the following thermochemical equations:

- a) formation of sulfur dioxide gas $\Delta\text{H} = -297 \text{ kJ/mol}$
b) decomposition of methane gas $\Delta\text{H} = +75 \text{ kJ/mol}$
c) combustion of propane gas ($\text{C}_3\text{H}_{8(\text{g})}$) $\Delta\text{H} = -2250. \text{ kJ/mol}$

Temperature-Heat Graph for a Pure Substance

States present	Solid	Solid + Liquid	Liquid	Liquid + Gas	Gas
Graph of Temperature vs Heat					
	What happens in this region?	Solid warms or cools	Solid melts or liquid freezes	Liquid warms or cools	Liquid vaporizes or gas condenses
Equation	$Q = mc\Delta T$	$Q = m * \Delta H$ $Q = n * \Delta H$	$Q = mc\Delta T$	$Q = m * \Delta H$ $Q = n * \Delta H$	$Q = mc\Delta T$

Specific heats for water:

solid $c = 2.1 \text{ J/g}^\circ\text{C}$
 liquid $c = 4.19 \text{ J/g}^\circ\text{C}$
 gas $c = 2.0 \text{ J/g}^\circ\text{C}$

ΔH values

fusion: 334 J/g or 6.01 kJ/mol
 freezing: -334 J/g or -6.01 kJ/mol
 condensation: -2.26 kJ/g or -40.7 kJ/mol
 vapourization: 2.26 kJ/g or 40.7 kJ/mol

ΔH refers to the change in heat content stored in the bonds that make up a molecule. Heat content is usually referred to as ENTHALPY.

Sample Problem:

Calculate the heat required to raise the temperature of 100.0 g of water from -1.0°C to $+1.0^\circ\text{C}$. The specific heat capacity of $\text{H}_2\text{O}_{(s)}$ is $2.06 \text{ J/g}^\circ\text{C}$ and of $\text{H}_2\text{O}_{(l)}$ is $4.19 \text{ J/g}^\circ\text{C}$. The molar heat of fusion is 6.01 kJ/mol .

• Specific Heat and Phase Changes Problems

1. The heat energy required to melt 1.00 g of ice at 0.0 °C is 333 J. If one ice cube has a mass of 62.0 g and a tray contains 20 ice cubes, how much energy is required to melt a tray of ice cubes?
2. How much heat energy would be required to raise the temperature of 454 g of Pb from room temperature (25.0 °C) to its melting point (327.0 °C) and then melt the lead at 327.0 °C? The specific heat of lead is 0.159 J/g·°C and the heat of fusion is 24.7 J/g.
3. Mercury with a freezing point of -39.0 °C, is the only metal that is a liquid at room temperature. How much heat would be released by Hg if 1.00 mL of it cools from 23.0 °C to -39.0 °C and then freezes to a solid? The density of Hg is 13.6 g/mL. Its specific heat is 0.138 J/g·°C and its heat of fusion is 11 J/g.
4. Liquid helium and liquid nitrogen are both used as coolants. He_(l) boils at 4.21 K and N_{2(l)} boils at 77.35 K. The specific heat of liquid helium near its boiling point is 4.25 J/g·K and the specific heat of liquid nitrogen near its boiling point is 1.95 J/g·K. Which liquid is a better coolant on a per-gram basis near its boiling point and which is better at its boiling point?
5. The hydrocarbon benzene, C₆H₆, boils at 80.1°C. How much heat energy is required to heat 1.00 kg of it from 20.0 °C to its boiling point and then change the liquid completely to a vapour? The specific heat capacity of liquid benzene is 1.74 J/g·°C and the heat of vaporization of benzene is 395 J/g.
6. a) Sketch a graph of the phase change involved in heating H₂O_(s) at 0.0 °C to the gas phase at 100.0 °C.

b) Calculate the quantity of heat required to convert four ice cubes (62.0 g each) from the solid phase at 0.0 °C to the gas phase at 100.0 °C. The heat of fusion is 333 J/g and the heat of vaporization is 2260. J/g.

• Supplemental Experiment: Energy of Combustion of Paraffin

Everyone knows the flame of a candle releases energy. In this experiment, you will determine how much heat is released.

What is the source of this energy? We find that energy is needed to break chemical bonds and energy is released when chemical bonds form. The main source of energy for the candle flame is found in the balance of these energy exchanges. When a candle burns, more energy is released when the bonds form to make products than is needed to break bonds in the reactants. The reactants are mainly candle wax and oxygen. The products are carbon dioxide, water, and soot.

The amount of heat produced will depend on the amount of wax burned. The procedure is to determine the temperature increase of a known mass of water by adding to it the heat produced by burning a known mass of candle wax.

Procedure:

1. Attach a candle to a piece of plastic petri dish.
2. Set up the apparatus as shown in Fig. 1. First, without the large can in place, adjust the height of the smaller can so that the bottom of the can is about 1 cm above the tip of the wick. The tip of the flame should almost but not quite reach the bottom of the can.
3. Mass the candle and base to the nearest 0.001 g. **Record the mass.** Place the candle and base in position inside the large can which protects the candle from drafts. (See Fig. 1.)
4. Using a graduated cylinder, measure the volume of tap water to fill the small can about two-thirds full. **Record the volume.**
5. Read and **record the temperature of the water to the nearest 0.5 °C.** Light the candle and quickly place the can of water in position. Heat the water, stirring it gently, until it reaches a temperature above room temperature. Carefully blow out the candle flame. Continue to stir the water, while watching the thermometer reading, until the highest temperature is reached. **Record the highest temperature to the nearest 0.5 C°.**
6. Mass the candle and base to the nearest 0.001 g. Make certain that any drippings from the candle are massed with it. **Record the mass.**

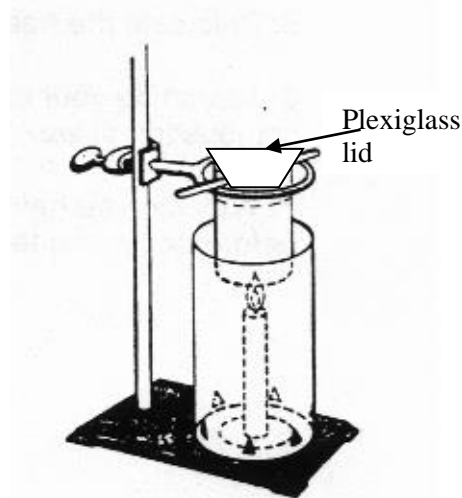


Figure 1

Data Table:

Initial mass of candle + dish	g
Final mass of candle + dish	g
Initial temperature of water	°C
Final temperature of water	°C
Volume of water in can	mL

Data Analysis:

1. Determine the mass of water heated (Density of water = 1.000 g/mL).
2. Determine the temperature change, ΔT , of the water.
3. Calculate the quantity of heat absorbed by the water in the can.
4. Determine the mass of candle wax burned.
5. Assuming your candle was a pure wax ($C_{25}H_{52(s)}$) calculate the number of moles of wax consumed to heat the water.
6. Calculate the molar heat of combustion of wax. $\Delta H_{\text{combustion}}$ in kJ/mol.

Conclusion:

1. The accepted value for the Heat of Combustion of Paraffin is 15380 kJ/mol. Determine the % error.
2. Explain any discrepancies.

• **TOPIC-4: PLENTY OF ENERGY**

A neutralization reaction is one in which an acid and a base neutralize each other producing a salt and water. We will now study quantitatively the heat involved in a neutralization reaction.

• **EXPERIMENT-4: Neutralization**

Purpose: To investigate heat transfer in a neutralization reaction.

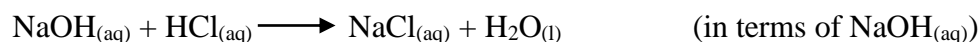
- Step- 1: Add 25.0 mL of 1.0 mol/L hydrochloric acid ($\text{HCl}_{(\text{aq})}$) into the calorimeter.
- Step-2: Record the room temperature. (Assume the temperature of the $\text{HCl}_{(\text{aq})}$ and $\text{NaOH}_{(\text{aq})}$ are the same as the room temperature.)
- Step-3: Add 25.0 mL of 1.0 mol/L sodium hydroxide ($\text{NaOH}_{(\text{aq})}$) into the calorimeter.
- Step-4: Immediately close the lid of the calorimeter Be sure the thermometer remains with the lid and *gently* stir to mix.
- Step-5: Reading the thermometer of the calorimeter, record the highest temperature reading. (Refer to as the final temperature, T_f of neutralization.)
- Step-6: Repeat steps 1 to 6 above then calculate and record the average value of ΔT .

$V_{\text{NaOH}(\text{aq})}$ (\pm mL)	$T_{i,\text{NaOH}(\text{aq})}$ (\pm °C)	$V_{\text{HCl}(\text{aq})}$ (\pm mL)	$T_{i,\text{HCl}(\text{aq})}$ (\pm °C)	T_i (\pm °C)	T_f (\pm °C)	$\Delta T = T_f - T_i$ (°C)
Average ΔT						

1) Using the heat formula, calculate the heat released in the neutralization reaction of your experiment. Use a value of 4.19 J/g.^oC for the specific heat capacity of water. (Note that we disregard the heat absorbed by the NaOH and HCl solutes.)

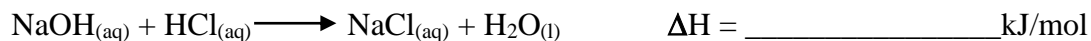
2) Determine the number of moles of hydrochloric acid and of sodium hydroxide that reacted in the neutralization. Hint: Molarity (M) = moles of solute per litre of solution (mol/L).

3) Calculate the "**molar heat of neutralization**" for the reaction.



Note: Your teacher (or lab technician) will write on the blackboard the average ΔH value obtained by the class as calculated above.

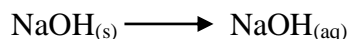
4) Below, fill-in the "class average" ΔH value for the neutralization reaction:



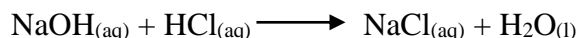
◆ **End of experiment**

• TOPIC-5: ADDING HEAT

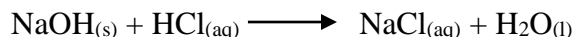
During experiment-3, we studied the heat involved when a substance dissolves in water. One of the solutes was "solid" sodium hydroxide (NaOH). This reaction is:



In experiment -4, we used an "aqueous" solution of sodium hydroxide to neutralize hydrochloric acid. This reaction is:



We measured the heat of reaction for both reactions separately. However, what would be the heat involved if we were to perform the reaction in a single step, that is, by adding solid sodium hydroxide directly into hydrochloric acid? This single step reaction would be similar to the above except that "solid" NaOH is used rather than "aqueous" NaOH. Like this:



• EXPERIMENT-5: Two in One

To answer the above question, we will perform an experiment.

Purpose: To find the heat of reaction ΔH for a neutralization reaction.

- Step-1: Mass 0.05 mol of sodium hydroxide ($\text{NaOH}_{(s)}$) into a calorimeter. Record mass.
- Step-2: Read and record the room temperature in the table below. (assume the temperature of $\text{HCl}_{(aq)}$ is the same as the room temperature.)
- Step-3: Add 50.0 mL of 1.0 mol/L hydrochloric acid ($\text{HCl}_{(aq)}$) into the calorimeter.
- Step-4: Cover with lid and stir *gently* with the thermometer.
- Step-5: Read and record the maximum temperature of the solution.
- Step-6: Repeat steps 1 to 5 above then calculate and record the average value of ΔT .

1) Record your experimental data in the table below:

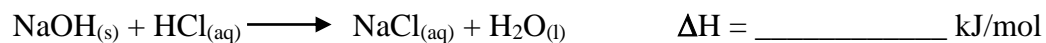
$m_{\text{NaOH(s)}}$ (\pm g)	$n_{\text{NaOH(s)}}$ (mol)	$V_{\text{HCl(aq)}}$ (\pm mL)	T_i (\pm °C)	T_f (\pm °C)	$\Delta T = T_f - T_i$ (°C)
Average ΔT					

2) Calculate the quantity of heat released during the reaction.

3) Determine the molar heat when sodium hydroxide ($\text{NaOH}_{(s)}$) is dissolved in $\text{HCl}_{(aq)}$.

Note: Your teacher (or lab technician) will write on the blackboard the average ΔH value obtained by the class as calculated above.

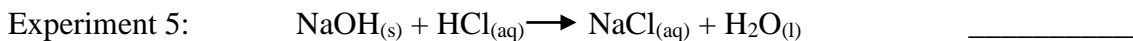
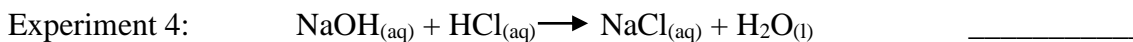
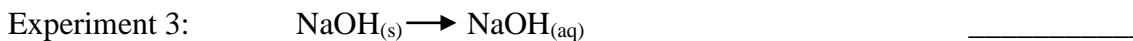
4) Below, fill-in the "class average" ΔH value of the neutralization reaction:



◆ **End of experiment**

Let's summarize the "class average" ΔH values below:

ΔH



From the above experimental values of ΔH , can you discover a law concerning the molar heats of reactions?

The Russian chemist Germaine Hess discovered experimentally that regardless of how many steps are involved in a chemical reaction, the heat flow (endo or exo) only depends on the initial and final states. Today we call his discovery **Hess's Law**:

Hess's Law of Constant Heat Summation: The change in energy in a chemical reaction is the **same** regardless of the path (steps) by which the reaction occurs. i.e. thermochemical equations can be added together like algebraic equations.

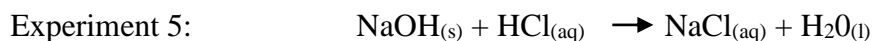
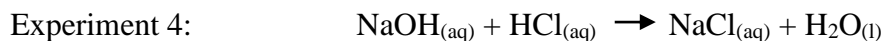
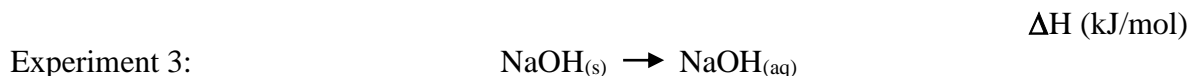
• STATE FUNCTIONS

A change in a system which is independent of the path (or steps) by which the change takes place (from the initial to the final state) is called a *state function*. Conversely, systems which do depend upon their path are called *path functions* such as time.

As an example of a state function (Hess' Law) consider height. Height is a state function. Thus, for example, in going up one floor, there are many paths which may be taken. Regardless of which path, however, in each case the end result is the same, we are one floor higher.

• TOPIC-6: REACTION STEPS

We studied the behavior of three energy systems experimentally. Continuing our study, we will use the standard values generally accepted by chemists for these same systems. The reactions were:



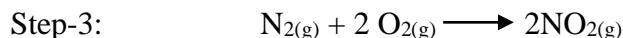
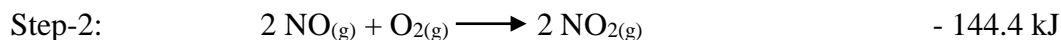
The equations above represent the same phenomenon viewed from different aspects.

The first equation represents the dissolving of sodium hydroxide, an exothermic reaction giving off heat. The second equation represents the reaction between the dissolved sodium hydroxide with hydrochloric acid, another exothermic reaction.

The third equation represents the same reaction as the first two combined. That is to say, as if the sodium hydroxide particles react directly with hydrochloric acid. In reality, of course, we know that the NaOH molecules must first dissolve and then react with the HCl molecules. It would seem logical, therefore, that the heat given off by the first two reactions should be the same as the heat given off by the third reaction.

This is exactly what Hess' Law predicts; "the quantity of heat involved in a chemical reaction is independent of the number of steps in the reaction; the quantity of heat is dependent only upon the starting and ending conditions."

Using the following equations, calculate the ΔH value for the chemical reaction involved when nitrogen ($\text{N}_{2(g)}$) and oxygen gas ($\text{O}_{2(g)}$) make nitrogen dioxide ($\text{NO}_{2(g)}$):



Molar heat of reaction: _____

In order to properly specify the molar heat of reaction without ambiguity, it is important to use a convention. The reason for this is because there are two ways to specify the molar heat of reaction. For example, in specifying the molar heat of reaction for ammonia, one method would be to specify the molar heat by writing the ΔH value to the right of the equation:



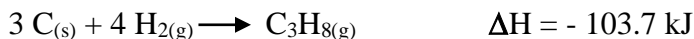
Using the above convention, there can be no confusion as the equation tells us that for every mole of $\text{N}_{2(\text{g})}$ reacting with three moles of $\text{H}_{2(\text{g})}$, we get two moles of ammonia formed. And, since the ΔH value is negative, we know that energy is lost by the reactants. That is, this is an exothermic reaction.

However, we can also specify the heat of reaction without the ΔH notation. In this case, the heat of reaction is specified as a species in the equation. That is, as a reactant or as a product. For the chemical equation above, we have three possibilities:

- ◆ 92.0 kJ per mole of $\text{N}_{2(\text{g})}$ -92.0 kJ/mol $\text{N}_{2(\text{g})}$
- ◆ 92.0 kJ per 3 moles of $\text{H}_{2(\text{g})}$ or 30.7 kJ per mole of $\text{H}_{2(\text{g})}$ -30.7 kJ/mol $\text{H}_{2(\text{g})}$
- ◆ 92.0 kJ per 2 moles of $\text{NH}_{3(\text{g})}$ or 46.0 kJ per mole of $\text{NH}_{3(\text{g})}$ -46.0 kJ/mol $\text{NH}_{3(\text{g})}$

The three possibilities above demonstrate that the expression kJ/mol can be ambiguous. To eliminate any ambiguity, we express the molar heat by **specifying a particular substance** in the reaction such as -46.0 kJ/mol $\text{NH}_{3(\text{g})}$.

Consider the following chemical equation and molar heat:



Specify the different ways in which the molar heat can be specified:

$$\Delta H = \underline{\hspace{10em}} \text{ kJ/mol C}_{(\text{s})}$$

$$\Delta H = \underline{\hspace{10em}} \text{ kJ/mol H}_{2(\text{g})}$$

$$\Delta H = \underline{\hspace{10em}} \text{ kJ/mol C}_3\text{H}_{8(\text{g})}$$

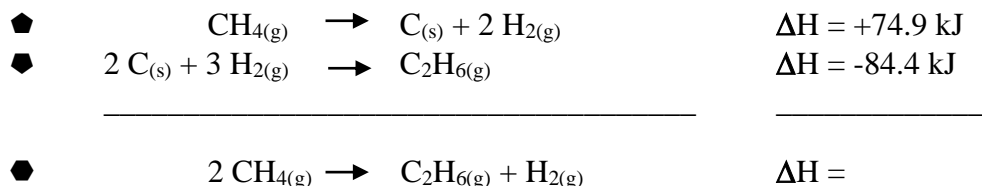
1. Knowing that $\Delta H = + 11.0 \text{ kJ/mol H}_{2(g)}$, find the heat of reaction for the equation:



2. The heat of reaction for a system is $-92.0 \text{ kJ/mol N}_{2(g)}$. Fill-in the ΔH value for this system:



3. To calculate the molar heat of reaction for the production of ethane (C_2H_6), we can begin with the reactions:



In solving for ΔH , it is important that the individual steps add up to the overall reaction. In the case above, we must check that the sum of the first two equations add up to the third equation.

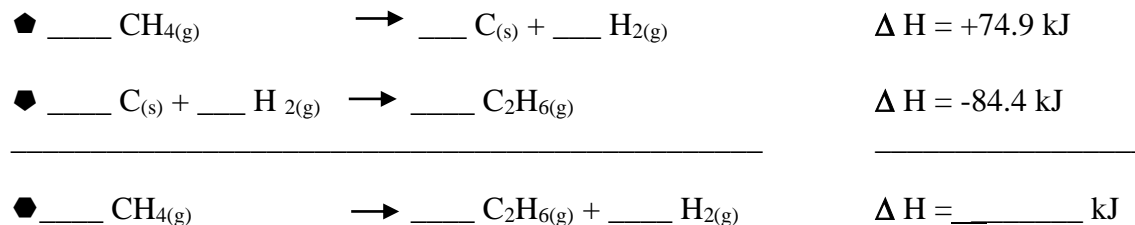
In the overall (or final) equation above, note that there are two moles of $\text{CH}_4(g)$.

How many times must equation-1 be used ? _____

What happens to the heat involved ? _____

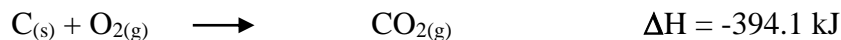
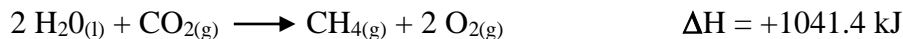
How many times must equation-2 be used? _____

Using the equations above fill-in the following coefficients in order to fulfill the above criteria:



Solve (algebraically) the first two reactions to obtain the third reaction.

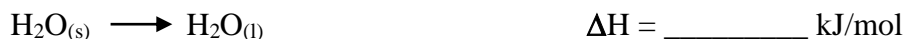
4. Applying Hess's law, find the heat of reaction for the following system:



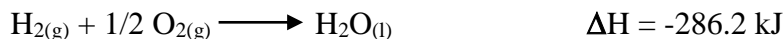
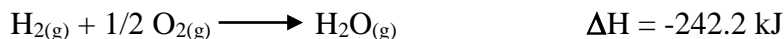
5. To make ice, we must remove 6.01 kJ of heat for every 18.02 g of water. Knowing this, find the heat of reaction for the following chemical equation:



What would be the value of ΔH for the reverse reaction?



6. Using the following equations, calculate the heat of reaction for the evaporation and condensation of water.



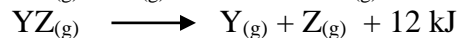
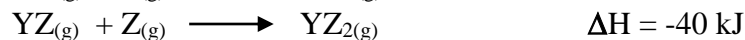
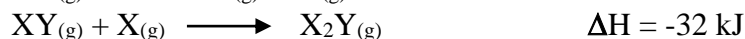
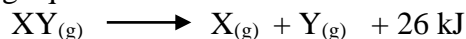
7. What is the quantity of heat involved when carbon monoxide converts into carbon dioxide?



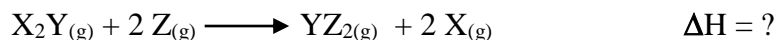
8. Natural gas (methane) transforms into propane by the following reaction. Determine the ΔH value for this reaction.



9. Given the following equations:



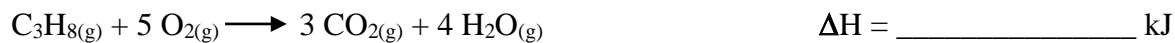
Determine ΔH for:



10. The following reaction represents the combustion of methane gas. Find the heat of combustion for this fuel.



The following equation represents the combustion of propane. Find the heat of combustion for propane.



In comparing methane with propane, which is a better fuel? Why?

11. The following reaction represents the combustion of ethylene. Find the heat of reaction for ethylene.



12. Determine the heat of reaction for the following reaction:

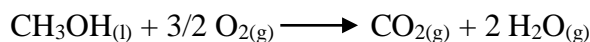


• **TOPIC-7: FOR THE PROS**

1. Suppose we required 200.0 mL of lukewarm water (40.0 °C). How could we obtain this water by mixing hot and cold water. Assume that 75.0 g of hot water at 75.0 °C is available.
2. Carbon disulfide reacts with oxygen to produce carbon dioxide and sulfur dioxide. Determine the heat of reaction:



3. The heating element of a "fondue" plate warmer consists of a self-contained bottle burning methanol (CH₃OH_(l)). When full the bottle holds 50.0 g of methanol. Find the quantity of heat released when one bottle of methanol completely burns.



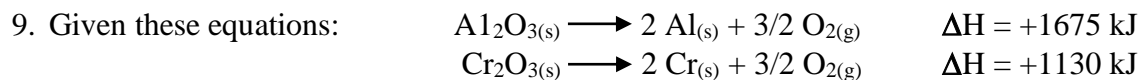
4. In a calorimeter, we burn 3.52 g of wax (C₂₅H_{52(s)}) thereby raising the water temperature by 18.5 °C. If the calorimeter holds 2 000.0 g of water, calculate the heat of combustion for wax.
5. During the combustion of 0.200 mol of a hydrocarbon, the temperature of 1500.0 g of water in a calorimeter rose from 18.4 °C to 56.3 °C. Determine the molar heat for this reaction.
6. The combustion of one gram of coal releases approximately 30.6 kJ of heat. What mass of coal must be burned in order to raise the temperature of 2 000.0 g of water from a temperature of 10.0 °C to 90.0 °C?
7. We want to freeze 1.00 kg of lukewarm water, which is at a temperature of 15.0 °C. We do this in two steps. During the first step, we cool the water from 15.0 °C to 0.0 °C. Next, we freeze the water according to the following equation:



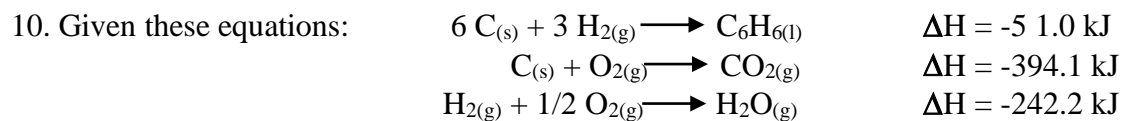
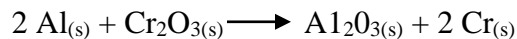
Calculate the quantity of heat involved for these two sequential processes.

Use the value 4.19 J/g·°C as the specific heat capacity of water for the range 15.0 °C to 0.0 °C.

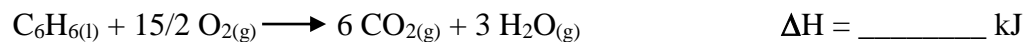
8. Calculate the heat required to raise the temperature of 100.0 g of water from -1.0 °C to + 3.0 °C. The specific heat capacity of H₂O_(s) is 2.06 J/g·°C and of H₂O_(l) is 4.19 J/g·°C. The molar heat of fusion is 6.01 kJ/mol.



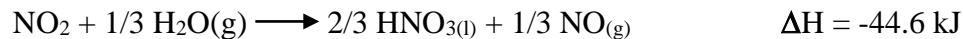
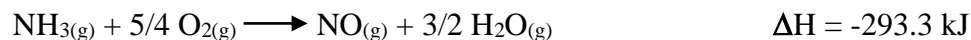
Calculate the heat of reaction for the following equation:



11. Determine the molar heat of combustion for benzene as per the reaction below:



12. Given the following equations, calculate the heat of combustion for ammonia, NH_3 .

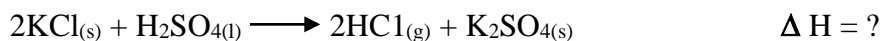


Problems: Hess's Law

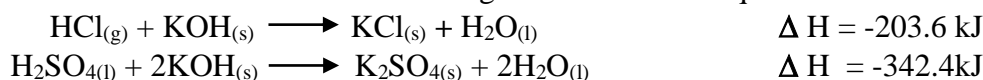
1. Use the first 2 equations below to find the value of ΔH for the last equation.



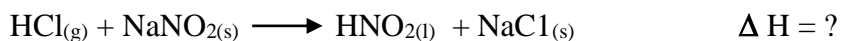
2. We can generate hydrogen chloride by heating a mixture of sulfuric acid and potassium chloride according to the reaction:



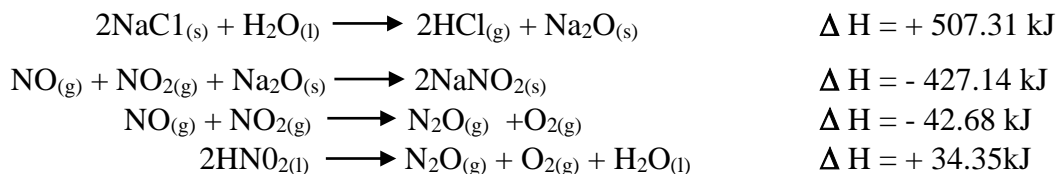
Calculate ΔH for this reaction from the following thermochemical equations.



3. Calculate ΔH for the following reaction, which describes the preparation of an unstable acid, $\text{HNO}_2(\text{l})$, nitrous acid.



Use the following thermochemical equations.



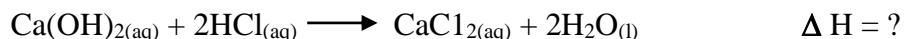
4. Copper metal can be obtained by roasting copper oxide, $\text{CuO}(\text{s})$, with carbon monoxide. Calculate the ΔH for this reaction:



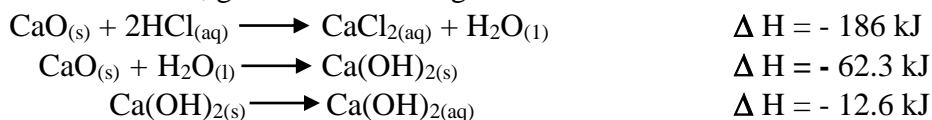
The following thermochemical data are known.



5. An aqueous solution of calcium hydroxide is called limewater. It neutralizes hydrochloric acid as follows:



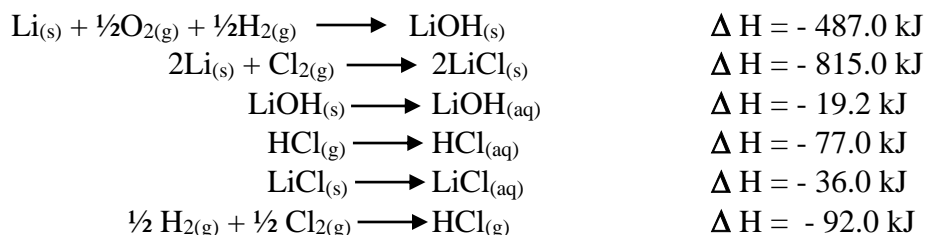
Calculate ΔH for this reaction, given the following thermochemical data.



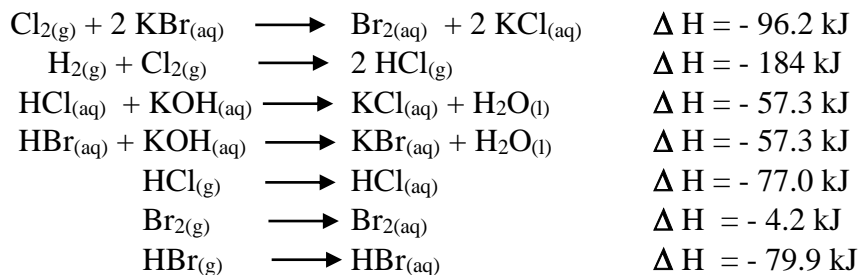
6. What is the value of ΔH for neutralization of lithium hydroxide by hydrochloric acid according to the following equation:



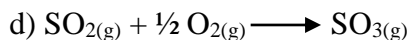
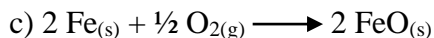
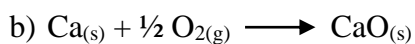
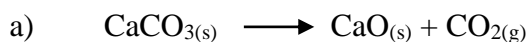
The following thermochemical data can be used. Also, you will need to write one (1) additional formation equation using data from the Heat of Formation Table (page 3-40).



7. (a) Write the thermochemical equation for the formation of 1 mole of $\text{HBr}_{(\text{g})}$
- (b) Use the following data to find the ΔH for the equation you wrote in (a).



8. Which of the following equations represent a formation (combination, synthesis) reaction ?



HEATS OF REACTION BETWEEN ELEMENTS
(REFERENCE TABLE)

ELEMENTS	→ COMPOUND	NAME	HEAT OF FORMATION (kJ/mole of product)
$\text{H}_{2(\text{g})} + 1/2 \text{O}_{2(\text{g})}$	→ $\text{H}_2\text{O}_{(\text{g})}$	Water vapour	-242.2
$\text{H}_{2(\text{g})} + 1/2 \text{O}_{2(\text{g})}$	→ $\text{H}_2\text{O}_{(\text{l})}$	Liquid water	-286.2
$\text{H}_{2(\text{g})} + 1/2 \text{O}_{2(\text{g})}$	→ $\text{H}_2\text{O}_{(\text{s})}$	Solid water (ice)	-292.2
$\text{C}_{(\text{s})} + 1/2 \text{O}_{2(\text{g})}$	→ $\text{CO}_{(\text{g})}$	Carbon monoxide	-110.7
$\text{C}_{(\text{s})} + \text{O}_{2(\text{g})}$	→ $\text{CO}_{2(\text{g})}$	Carbon dioxide	-394.1
$\text{C}_{(\text{s})} + 2 \text{H}_{2(\text{g})}$	→ $\text{CH}_{4(\text{g})}$	Methane	-74.9
$2 \text{C}_{(\text{s})} + 2 \text{H}_{2(\text{g})}$	→ $\text{C}_2\text{H}_{4(\text{g})}$	Ethylene	+52.3
$2 \text{C}_{(\text{s})} + 3 \text{H}_{2(\text{g})}$	→ $\text{C}_2\text{H}_{6(\text{g})}$	Ethane	-84.4
$3 \text{C}_{(\text{s})} + 4 \text{H}_{2(\text{g})}$	→ $\text{C}_3\text{H}_{8(\text{g})}$	Propane	-103.7
$\text{C}_{(\text{s})} + 2 \text{H}_{2(\text{g})} + 1/2 \text{O}_{2(\text{g})}$	→ $\text{CH}_3\text{OH}_{(\text{l})}$	Methanol	-238.8
$\text{C}_{(\text{s})} + 1/4 \text{S}_{8(\text{s})}$	→ $\text{CS}_{2(\text{l})}$	Carbon disulfide	+89.8
$1/8 \text{S}_{8(\text{s})} + \text{O}_{2(\text{g})}$	→ $\text{SO}_{2(\text{g})}$	Sulfur dioxide	-297.3
$1/8 \text{S}_{8(\text{s})} + 3/2 \text{O}_{2(\text{g})}$	→ $\text{SO}_{3(\text{g})}$	Sulfur trioxide	-396.3
$\text{H}_{2(\text{g})} + 1/8 \text{S}_{8(\text{s})} + 2 \text{O}_{2(\text{g})}$	→ $\text{H}_2\text{SO}_{4(\text{l})}$	Sulfuric acid	-810.9
$1/2 \text{N}_{2(\text{g})} + 1/2 \text{O}_{2(\text{g})}$	→ $\text{NO}_{(\text{g})}$	Nitric oxide	+90.3
$1/2 \text{N}_{2(\text{g})} + \text{O}_{2(\text{g})}$	→ $\text{NO}_{2(\text{g})}$	Nitrogen dioxide	+33.9
$\text{N}_{2(\text{g})} + 3/2 \text{O}_{2(\text{g})}$	→ $\text{N}_2\text{O}_{3(\text{g})}$	Dinitrogen trioxide	+83.8
$\text{N}_{2(\text{g})} + 2 \text{O}_{2(\text{g})}$	→ $\text{N}_2\text{O}_{4(\text{g})}$	Dinitrogen tetroxide	+83.8
$1/2 \text{N}_{2(\text{g})} + 3/2 \text{H}_{2(\text{g})}$	→ $\text{NH}_{3(\text{g})}$	Ammonia	-46.0
$2 \text{C}_{(\text{s})} + 3 \text{H}_{2(\text{g})}$	→ $\text{C}_2\text{H}_{6(\text{g})}$	Ethane	-84.4
$3 \text{C}_{(\text{s})} + 4 \text{H}_{2(\text{g})}$	→ $\text{C}_3\text{H}_{8(\text{g})}$	Propane	-103.7
$1/2 \text{H}_{2(\text{g})} + 1/2 \text{I}_{2(\text{s})}$	→ $\text{HI}_{(\text{g})}$	Hydrogen iodide	+26.0
$1/2 \text{H}_{2(\text{g})} + 1/2 \text{I}_{2(\text{g})}$	→ $\text{HI}_{(\text{g})}$	Hydrogen iodide	-5.0
$2 \text{Fe}_{(\text{s})} + 3/2 \text{O}_{2(\text{g})}$	→ $\text{Fe}_2\text{O}_{3(\text{s})}$	Iron(III) oxide	-197.0
$3 \text{Fe}_{(\text{s})} + 2 \text{O}_{2(\text{g})}$	→ $\text{Fe}_3\text{O}_{4(\text{s})}$	Iron(II,III) oxide	-1120.0
$2 \text{Al}_{(\text{s})} + 3/2 \text{O}_{2(\text{g})}$	→ $\text{Al}_2\text{O}_{3(\text{s})}$	Aluminum oxide	-1668
$\text{Cu}_{(\text{s})} + 1/2 \text{O}_{2(\text{g})}$	→ $\text{CuO}_{(\text{s})}$	Copper(II) oxide	-155.0
$\text{Mg}_{(\text{s})} + 1/2 \text{O}_{2(\text{g})}$	→ $\text{MgO}_{(\text{s})}$	Magnesium oxide	-601.1
$\text{Na}_{(\text{s})} + 1/2 \text{Cl}_{2(\text{g})}$	→ $\text{NaCl}_{(\text{s})}$	Sodium chloride	-410.5

Experiment: Molar Heat of Fusion of Ice (ΔH_{fusion})

Purpose: To determine the amount of heat energy needed to melt one mole of ice using calorimetry.

Hypothesis: Assuming $-Q_{\text{water}} = +Q_{\text{ice}}$, the heat lost by the water as it cools can be calculated by (formula).

/3

The mass of the ice can be determined by (procedure).

The molar heat of fusion can be calculated by (formula) and will be approximately the accepted value of +6.01 kJ/mol.

Materials: (List with bullet points)

/2

Procedure: (Number the steps.)
(Three trials required.)

/3

Diagram: (Labelled.)

/2

Data: (In a table, include all measurements and uncertainty values)

/2

Analysis: (All calculations using significant figures, experimental error (E) and % error using the accepted value of +6.01 kJ/mol)

/5

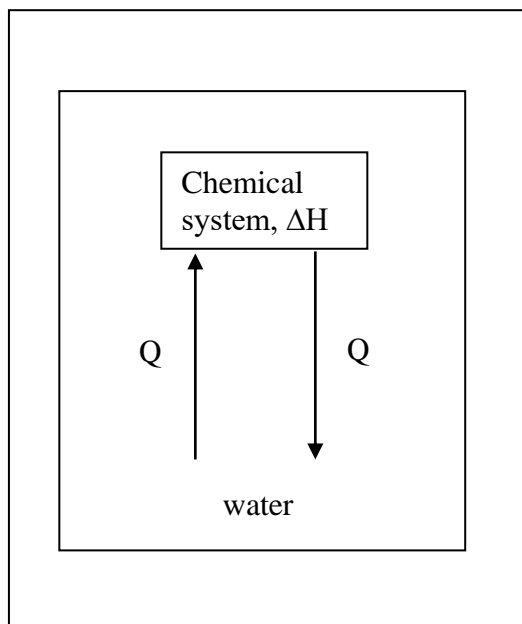
Conclusion: The molar heat of fusion was determined to be _____.

/2

Calorimetry **is/is not** an effective method because the source of error was **less/greater** than 10%.

Error Analysis: (One source of experimental error – how it affected the calculated value of ΔH_{fusion} – recommendation on how to reduce or eliminate this error)

/2



Calorimetry is the technological process of measuring energy changes using an insulated system called a calorimeter.

The system inside the calorimeter undergoes either a phase change or a chemical change.

If heat energy is absorbed by the system, ΔH is positive, the reaction is endothermic and the temperature of the water decreases.

If the heat energy is released by the system, ΔH is negative, the reaction is exothermic and the temperature of the water increases.

1. What are 2 assumptions made when using a calorimeter to measure heat exchange?
2. In an experiment, 10.0 g of urea is dissolved in 150.0 mL of water in a calorimeter. The temperature of the water changes from 20.4 to 16.7 °C. Calculate the molar heat of solution of urea, NH_2CONH_2 .
3. When 0.187 g of benzene, C_6H_6 , is burned in a bomb calorimeter, the surrounding water bath rises in temperature by 7.48 °C. If the bath contains 250.0 g of water, what is the molar heat of combustion of benzene?
4. When 1.045 g of CaO is added to 50.0 mL of water at 25.0 °C in a calorimeter, the temperature of the water increases to 32.3 °C. Calculate the molar heat of solution.

5. A student pours some liquid paraffin, $C_{25}H_{52}$, at its melting point into a calorimeter and records the final temperature when the wax just solidifies. Use the data below to calculate the molar heat of solidification of paraffin wax.

Volume of water in calorimeter = 150.0 mL

Mass of paraffin per trial = 25.00 g

Trial	1	2	3
Final temperature ($^{\circ}C$)	27.1	27.7	27.5
Initial temperature ($^{\circ}C$)	20.4	21.2	20.9

6. A 10.0 g sample of gallium metal is added to 50.0 mL of water in a calorimeter. The temperature of the water changes from 27.8 to 24.0 $^{\circ}C$ as the gallium melts. Calculate the molar heat of fusion of gallium.

7. When 8.00 g of NaOH is added to a calorimeter containing 75.0 ml of a solution of HCl, the temperature of the solution increases from 25.0 to 33.5 $^{\circ}C$. Calculate the molar heat of neutralization for the NaOH. (Note: Assume the density of the solution is 1.00 g/mL.)

8. In a calorimeter, 50.0 mL of 2.0 M HNO_3 (nitric acid) is added to 50.0 mL of 2.0 M KOH. The temperature of the solution increases by 15.0 $^{\circ}C$. Calculate the molar heat of solution for the KOH. (Note: Assume the density of the solution is 1.00 g/mL.)

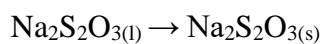
A student placed 20.0 g of liquid sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$, into 100.0 mL of cold water in a calorimeter. The following data was recorded.

Time (min.)	Temperature of H_2O ($^{\circ}\text{C}$)	Temperature of $\text{Na}_2\text{S}_2\text{O}_3$ ($^{\circ}\text{C}$)
0	14.6	52.0
1	17.6	50.0
2	19.6	48.2
3	21.6	48.2
4	22.0	48.2
5	23.0	46.6
6	23.4	44.0
7	23.6	40.0
8	23.8	32.0

1. Draw a graph of Temperature ($^{\circ}\text{C}$) vs Time (min.). Plot the data for the water and for the $\text{Na}_2\text{S}_2\text{O}_3$ on the same graph. Label each line.
2. What is the freezing point of $\text{Na}_2\text{S}_2\text{O}_3$? _____
3. Calculate ΔT for the water during the time when the $\text{Na}_2\text{S}_2\text{O}_3$ is freezing? _____

4. Using the ΔT from question 3 to calculate the amount of heat energy (Q) absorbed by the water during the time when the $\text{Na}_2\text{S}_2\text{O}_3$ is freezing.
5. Calculate the molar heat of freezing (ΔH) of $\text{Na}_2\text{S}_2\text{O}_3$.
6. Is the freezing of liquid $\text{Na}_2\text{S}_2\text{O}_3$ endothermic or exothermic? _____

7. What happens to the molecules of the liquid $\text{Na}_2\text{S}_2\text{O}_3$ during the freezing process?
8. Insert the kJ of heat energy from question 5 on the proper side of the equation below.



Hess' Law Problems - Workbook Answers

Chapter 3

Starting on p. 3-30

- | | |
|---|---|
| 33. -34.6 kJ/mol C _(s) ; -25.9 kJ/mol H _{2(g)} ; -103.7 kJ/mol C ₃ H _{8(g)} | 43. -803.6 kJ |
| 34. 33.0 kJ | 44. -2047.4 kJ |
| 35. -46.0 kJ | 45. propane |
| 36. 65.4 kJ | 46. -1324.9 kJ |
| 37. -74.9 kJ | 47. -1649.0 kJ |
| 38. -6.01 kJ/mol ; + 6.01 kJ/mol | 48. 125.0 g |
| 39. 44.0 kJ/mol ; -44.0 kJ/mol | 49. -1078.5 kJ |
| 40. -283.4 kJ | 50. -639.7 kJ/mol → Answer: Q = -998 kJ |
| 41. 121.0 kJ | 58. -545 kJ |
| 42. -22 kJ | 59. -3040.2 kJ |
| | 60. -4760.4 kJ |

Starting on p. 3-41

- | | |
|--------------|-------------|
| 1. -171.0 kJ | 5. -111 kJ |
| 2. 64.8 kJ | 6. -54.5 kJ |
| 3. -78.60 kJ | 7. -43 kJ |
| 4. -127.8 kJ | 8. b, c, d |

Starting on p. 3-47

- | | |
|----------------|-----------|
| 1. 35000 kJ | 5. 461 kJ |
| 2. 22100 kJ | 6. 516 kJ |
| 3. -720 kJ | 7. 30.°C |
| 4. 5.19 J/g·°C | |

Starting on p. 3-37 and 3-38

(Specific Heat Problems)

1. $Q_{\text{H}_2\text{O}} = 754 \text{ J}$; $Q_{\text{Al}} = 647 \text{ J}$; water
2. 101 kJ
3. a) water
b) $Q_{\text{H}_2\text{O}} = 1.49 \times 10^3 \text{ kJ}$
 $Q_{\text{EG}} = 953 \text{ kJ}$
4. 1.3 kJ
5. 335°C
6. 6.2°C
7. Li \rightarrow 24.8 J/mol•°C
Na \rightarrow 28.3 J/mol•°C
F₂ \rightarrow 31.3 J/mol•°C
Cl₂ \rightarrow 33.9 J/mol•°C

Starting on p. 3-38 and 3-39

(Specific Heat & Phase Change Problems)

1. 413 kJ
2. 33.0 kJ
3. - 0.27 kJ
4. Helium
5. 500. kJ
- 6b. 747 kJ

Starting on p. 3-35 and 3-36

(Calorimetry Problems)

51. $-1.55 \times 10^4 \text{ kJ/mol}$
52. $-1.19 \times 10^3 \text{ kJ/mol}$
53. 21.9 g
54. -397 kJ
55. 33.9 kJ
56. 0.84 kJ

Starting on p. 3-52 and 3-53

(Calorimetry Problems)

2. 14 kJ/mol
3. $-3.28 \times 10^3 \text{ kJ/mol}$
4. -82 kJ/mol
5. -59 kJ/mol
6. 5.6 kJ/mol
7. -13 kJ/mol
8. -63 kJ/mol