

# 7—THERMOCHEMISTRY: HEAT OF REACTION



Name: \_\_\_\_\_

Date: \_\_\_\_\_

Section: \_\_\_\_\_

## Objectives

- Measure the enthalpy of reaction for the decomposition of hydrogen peroxide
- Measure the heat capacity of a Styrofoam cup calorimeter using the heat of neutralization of a strong acid with a strong base
- Graph your temperature vs time data to find temperature change when solutions are mixed

## Pre-Laboratory Requirements

- Read chapter 6 in Silberberg
- Pre-lab questions (if required by your instructor)
- Laboratory notebook—prepared before lab (if required by your instructor)

## Safety Notes

- Eye protection must be worn at all times.
- Hydrochloric acid and sodium hydroxide are caustic and should not come in contact with your skin or clothing. Wear gloves when handling these chemicals. A lab coat or lab apron is recommended.
- The hydrogen peroxide used in this experiment is often used as an antiseptic. However, avoid unnecessary exposure to your skin and clothing. Do not splash hydrogen peroxide in your eyes.

## Discussion

When a chemical reaction occurs, bonds are broken and new bonds are formed. The breaking of bonds requires energy while the formation of new bonds releases energy. For most spontaneous chemical reactions, the energy released by the new bond formation is greater than that required to break the old bonds. In this process, energy is released (usually as heat) into the environment. These reactions are called **exothermic** and a minus sign is put in front of the quantity of energy **released**. Some reactions require energy from the environment to proceed; these are called **endothermic** reactions. Because energy must flow from the environment into the reactants for the reaction to occur, a plus sign is put in front of the amount of energy **absorbed**. The amount of energy released or absorbed in the reaction is called the **Heat of Reaction**,  $H_{\text{rxn}}$ .

If  $\Delta H_{\text{rxn}} = -145.3$  kJ/mole for a specific reaction, this reaction is exothermic, the energy released by new bond formation exceeds that required to break the old bonds, and for each mole of reaction 145.3 kJ are released. If a reaction gives off heat (or takes in heat), the heat has to go (or come from) somewhere—remember the law of conservation of energy. Adding heat to a system will increase the kinetic energy of the system. As a result of the added energy, the atoms will move faster and the temperature will increase. Temperature is a measure of how fast the atoms and molecules are moving, and can be changed by adding or removing heat energy. Every substance has its own unique response to added heat energy. For some materials, even a small amount of added heat will cause a large increase in temperature. For others materials, it takes a great deal of heat energy to raise the temperature even slightly. The ratio between the heat energy change and the temperature change is called the **heat capacity**.

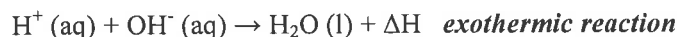
$$\text{Heat Capacity}(C) = \frac{\text{heat gained or lost}}{\text{change in temperature}} = \frac{q}{\Delta T} \quad (1)$$

The most common way to measure  $\Delta H_{\text{rxn}}$  is with a calorimeter. In its simplest form, a calorimeter is a container that has material to absorb heat released by a reaction, but does not lose heat to the environment. A Styrofoam cup containing water can be used as a calorimeter.

## Part I. Determination of the Heat Capacity of the Calorimeter

Let's look at what happens when the reaction between HCl and NaOH is carried out in a calorimeter. Beginning with the two aqueous solutions, the HCl solution is put in the calorimeter and the NaOH is put into another cup. Since HCl is a strong electrolyte, it is completely dissociated into  $\text{H}^+$  and  $\text{Cl}^-$  ions. NaOH is also a strong electrolyte and is completely dissociated into  $\text{Na}^+$  and  $\text{OH}^-$  ions in solution. If the solutions sit, they will become the same temperature as the surroundings (*i.e.* the room).

Once the reactants have equilibrated to an initial temperature ( $T_i$ ), the NaOH solution is poured into the calorimeter.  $\text{Na}^+$  and  $\text{Cl}^-$  are spectator ions so they don't participate in the reaction.  $\text{H}^+$  and  $\text{OH}^-$  are an acid and a base, respectively, and will react to form water.



In this reaction, an O-H bond is formed. Heat energy will be released by the formation of this bond. The heat energy released will heat the water and calorimeter. The temperature of the water and calorimeter will rise because it is accepting the heat given off by the reaction. In other words, the heat *released* by the reaction ( $q_{\text{rxn}}$ ) is *gained* by the water and calorimeter ( $q_{\text{cal}}$ ). Assuming that no heat is lost from the calorimeter (*i.e.* that this is a closed system), this heat exchange can be represented as:

$$\text{heat released by reaction} + \text{heat absorbed by calorimeter and water} = 0$$

$$q_{\text{rxn}} + q_{\text{cal}} = 0 \quad (2)$$

Remember, if heat is given off,  $q < 0$  and if heat is absorbed  $q > 0$ .

After the reaction has gone to completion and the heat has been completely absorbed by the calorimeter system (calorimeter + water), the calorimeter temperature will reach a maximum value ( $T_f$ ). This temperature will be higher than the initial temperature of the water. This change in temperature is denoted as  $\Delta T = T_f - T_i$ . This temperature change can be measured.

Since no container is perfectly insulating (*i.e.* heat leak proof) some of the heat will leak out of the container and the temperature of the calorimeter system will slowly fall. Eventually, all of the heat will leak out and the temperature will return to its initial value.

## Procedure

### Part I. Determination of the Heat Capacity of the Calorimeter

The calorimeter consists of two Styrofoam cups, a stir bar, 50 mL of water (formed from the reaction of HCl and NaOH), a lid, and a thermometer. All of these will absorb heat energy to raise their temperature. Since everything is in contact, all the components will have the same temperature after a very short time. The system is not perfect and leaks heat over time. You will determine how much heat energy it takes to raise the temperature of this system by 1 °C. To do this, you will carry out an experiment in which a known amount of heat is added to our

system, and measure the temperature change that results. The source of the added heat will be the chemical reaction between HCl and NaOH.

1. Obtain 3 cups, two lids, a stir-plate (found on the hot plate) and a stir bar, two 50.0 mL graduated cylinders, and a digital thermometer.
2. Nest two of the cups in each other (this is your calorimeter) and place the stir bar in the top cup.
3. Put the calorimeter on the stir plate, and add exactly 25.0 mL of 1.0 M HCl.
4. Start the stirrer at a moderate speed and place a lid on the calorimeter.
5. Add 25.0 mL of 1.0 M NaOH to a separate cup, and place a lid on this cup.
6. Measure and record the temperature of the solutions in the two cups every 30 seconds for 2 minutes
7. **Be sure to rinse the temperature probe with water and dry it when going from one solution to the other.**
8. At the end of the 2-minute monitoring period, quickly pour the NaOH into the cup containing the HCl and replace the lid on the calorimeter cup.
9. Record the temperature of the combined solutions every 30 seconds for an additional 9 minutes.
10. You will record temperature a total of 11 minutes (2 minutes before combining and 9 minutes after they were combined).
11. After 11 minutes, empty the water from the cups, dry them carefully, and repeat the experiment.

Data Table – Heat Capacity of the Calorimeter - Trial 1

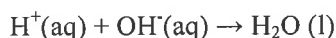
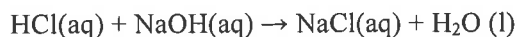
Time (min)	Temperature (°C)	Time (min)	Temperature (°C)	Time (min)	Temperature (°C)
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Data Table – Heat Capacity of the Calorimeter - Trial 2

Time (min)	Temperature (°C)	Time (min)	Temperature (°C)	Time (min)	Temperature (°C)
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## Calculations

First find the heat released by the reaction of NaOH and HCl. The balanced equations for the full and net ionic reactions are:



The acid-base neutralization reaction has been studied and is known to have a  $\Delta H_{\text{rxn}} = -55.84 \text{ kJ/mol}$ . This means that 55.84 kJ of energy are released for *each mole of water that is formed*. Calculate the moles of water formed during the reaction given the volumes and molarities of reactants used and then determine the amount of heat released by the reaction,  $q_{\text{rxn}}$ .

Assuming all of the heat released by the chemical reaction is absorbed by the calorimeter system, calculate  $q_{\text{cal}}$ .

The heat absorbed by the calorimeter system,  $q_{\text{cal}}$ , is equal to the heat capacity of the calorimeter system multiplied by the temperature change.

$$q_{\text{cal}} = C_{\text{cal}} \times \Delta T \quad (3)$$

To find  $\Delta T$ , make a plot of T vs. time for both trials. The plots should look similar to *Figure 1*. Graph paper is provided at the end of this experiment to use in preparing your graph.

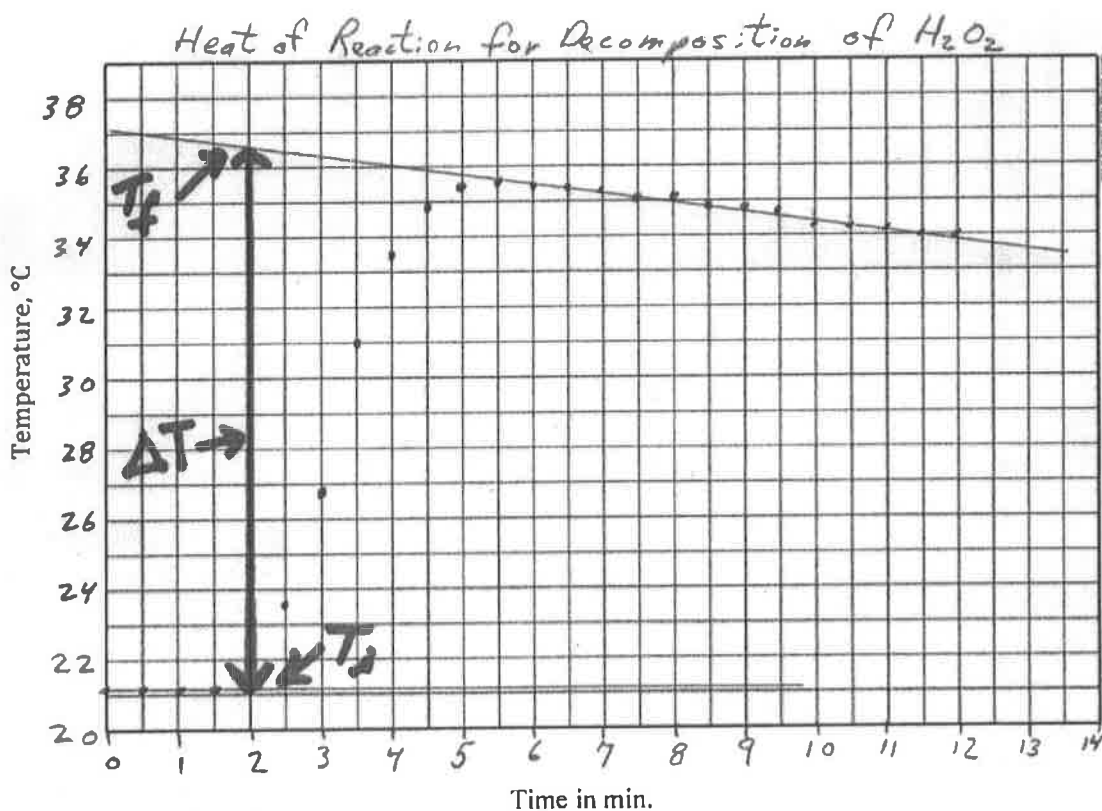


Figure 1. Graph of temperature data for the decomposition of H<sub>2</sub>O<sub>2</sub>. Note that the temperatures of the individual solutions, before mixing, were averaged and then plotted on the graph paper.

To calculate  $\Delta T$ , you need  $T_i$  (the initial temperature), and  $T_f$  (the final temperature) after the reaction has occurred, **but before any heat has had a chance to leak out**. To find  $T_i$ , average the two initial temperatures for the HCl and NaOH solutions immediately before the reaction occurred. To find the final temperature,  $T_f$ , in the absence of heat leakage, use the graph and extrapolate back from the temperatures after mixing to the time of mixing.  $\Delta T = T_f - T_i$ .

Determine the heat capacity of the calorimeter,  $C_{\text{cal}}$ . Find the average  $C_{\text{cal}}$ .

## Part II. Determination of the Heat of Reaction, $\Delta H_{rxn}$

In this part of the experiment you will determine the heat of reaction for the decomposition of hydrogen peroxide. Hydrogen peroxide interacts with a catalyst that results in the decomposition of the hydrogen peroxide into water and oxygen:



Here the  $\text{Fe}^{3+}$  ion acts as a catalyst to facilitate the reaction, but it is not transformed by the reaction. Also, note that 1 mole of water and  $\frac{1}{2}$  mole of oxygen is produced for each mole of  $\text{H}_2\text{O}_2$  that decomposes.

1. Rinse and dry the cups used in the calibration above and set up your calorimeter again
2. Transfer 40.0 mL of  $\text{H}_2\text{O}_2$  into the calorimeter (Be sure to record the **molarity** of the solution, which can be found on the bottle in the fume hood)
3. Transfer 10.0 mL of 0.5 M  $\text{Fe}(\text{NO}_3)_3$  to the single, separate cup
4. Begin stirring the solution in the calorimeter, position the lid, and take temperature readings on each solution every 30 seconds for 2 minutes
5. After 2 minutes, pour the 0.5 M  $\text{Fe}(\text{NO}_3)_3$  solution into the  $\text{H}_2\text{O}_2$  solution, replace the lid, and begin taking temperature measurements every 30 seconds for 10 minutes.
6. You will record temperature a total of 12 minutes (2 minutes before combining and 10 minutes after they were combined).
7. Empty the solution from the cups, rinse the cups with water, dry them, and repeat the experiment.

Data Table – Heat of Reaction for the Decomposition of  $\text{H}_2\text{O}_2$  - Trial 1

Time (min)	Temperature ( $^{\circ}\text{C}$ )	Time (min)	Temperature ( $^{\circ}\text{C}$ )	Time (min)	Temperature ( $^{\circ}\text{C}$ )
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Data Table – Heat of Reaction for the Decomposition of  $\text{H}_2\text{O}_2$  - Trial 2

Time (min)	Temperature ( $^{\circ}\text{C}$ )	Time (min)	Temperature ( $^{\circ}\text{C}$ )	Time (min)	Temperature ( $^{\circ}\text{C}$ )
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## Calculations

1. Once again, make a plot of T vs. Time and Calculate  $\Delta T$  for the two trials as before. The initial temperature is the average temperature of the two solutions. The final temperature is found by extrapolation from the graph.
2. Using the average  $C_{\text{cal}}$  and  $\Delta T$ , calculate  $q_{\text{cal}}$  and then find  $q_{\text{rxn}}$ . (Hint: Remember that the heat gained by the calorimeter was released by the reaction).
3. The calculated  $q_{\text{rxn}}$  is for the amount of material used in our trials. We want to determine the heat of reaction **per mole of  $\text{H}_2\text{O}_2$  decomposed**. To obtain this, determine how many moles of  $\text{H}_2\text{O}_2$  actually decomposed from the volumes and molarities of reactants used. Calculate the heat of reaction per one mole of  $\text{H}_2\text{O}_2$ :

$$\Delta H_{\text{rxn}} = q_{\text{rxn}} / \text{moles } \text{H}_2\text{O}_2.$$

4. Find the average  $\Delta H_{\text{rxn}}$ .
5. The literature value for  $\Delta H_{\text{H}_2\text{O}_2}$  is -94.6 kJ/mol. Calculate the error and percent relative error for your  $\Delta H_{\text{H}_2\text{O}_2}$ .

## Reference

Adapted from: Marzacco, C. J. "The Enthalpy of Decomposition of Hydrogen Peroxide: A General Chemistry Calorimetry Experiment" *J. Chem. Ed.* **76** 1517 (1999).

### Calculation Guide—Heat Capacity of Calorimeter

Volume of HCl used (mL): \_\_\_\_\_ Concentration of HCl used (M): \_\_\_\_\_

Moles of HCl used (mol): \_\_\_\_\_

Volume of NaOH used (mL): \_\_\_\_\_ Concentration of NaOH used (M): \_\_\_\_\_

Moles of NaOH used (mol): \_\_\_\_\_

Moles of water formed (mol): \_\_\_\_\_

From your graph, trial 1:  $T_i$  \_\_\_\_\_ °C  $T_f$  \_\_\_\_\_ °C  $\Delta T$  \_\_\_\_\_ °C

From your graph, trial 2:  $T_i$  \_\_\_\_\_ °C  $T_f$  \_\_\_\_\_ °C  $\Delta T$  \_\_\_\_\_ °C

$\Delta H_{f(H_2O)} = -55.84 \text{ kJ/mol}$

$q_{rxn}$  (kJ), trial 1 : \_\_\_\_\_  $q_{rxn}$  (kJ), trial 2: \_\_\_\_\_

$q_{cal}$  (kJ), trial 1 : \_\_\_\_\_  $q_{cal}$  (kJ), trial 2: \_\_\_\_\_  $q_{cal}$  (kJ), average: \_\_\_\_\_

$C_{cal}$  (J/°C): \_\_\_\_\_ Note: This is the average from 2 trials

### Calculation Guide—Heat of Reaction for Decomposition of $H_2O_2$

Volume of  $H_2O_2$  used (mL): \_\_\_\_\_ Concentration of  $H_2O_2$  used (M): \_\_\_\_\_

Moles of  $H_2O_2$  used (mol): \_\_\_\_\_

From your graph, trial 1:  $T_i$  \_\_\_\_\_ °C  $T_f$  \_\_\_\_\_ °C  $\Delta T$  \_\_\_\_\_ °C

From your graph, trial 2:  $T_i$  \_\_\_\_\_ °C  $T_f$  \_\_\_\_\_ °C  $\Delta T$  \_\_\_\_\_ °C

$C_p$ -Calorimeter (J/°C): \_\_\_\_\_ (From above)

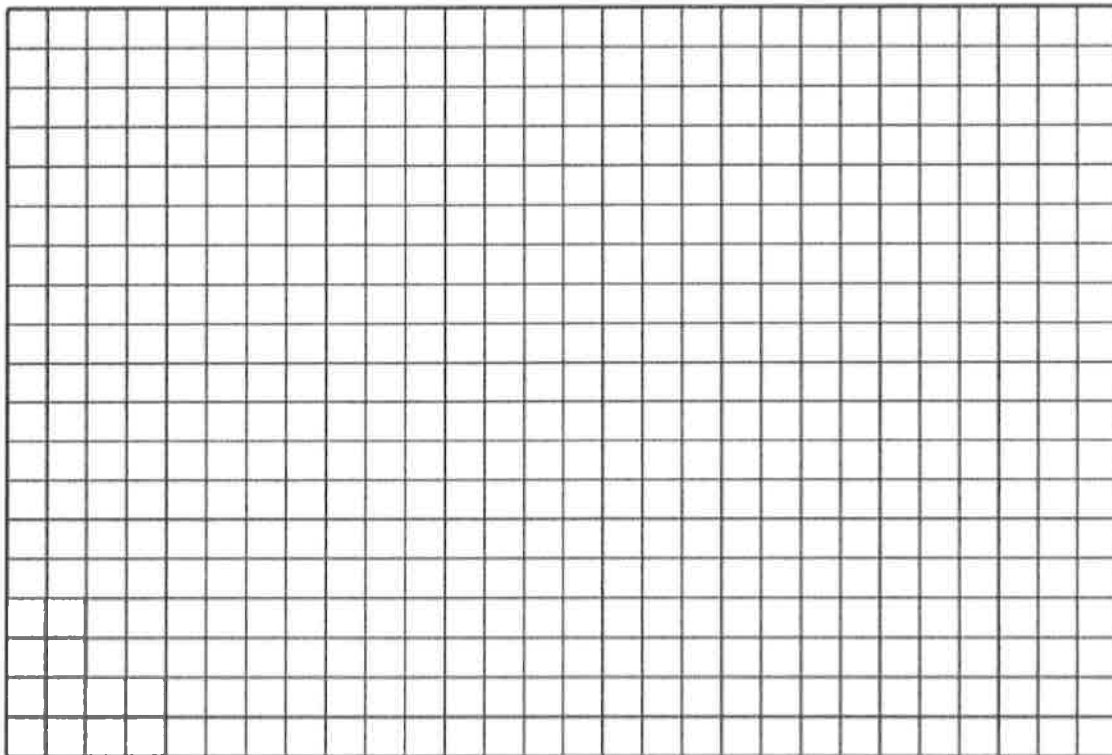
$q_{cal}$  (kJ), trial 1 : \_\_\_\_\_  $q_{cal}$  (kJ), trial 2: \_\_\_\_\_

$q_{rxn}$ (kJ), trial 1 : \_\_\_\_\_  $q_{rxn}$  (kJ), trial 2: \_\_\_\_\_  $q_{rxn}$  (kJ), average: \_\_\_\_\_

$\Delta H_{H_2O_2}$  (KJ/mol): \_\_\_\_\_ Error: \_\_\_\_\_ Percent relative Error: \_\_\_\_\_

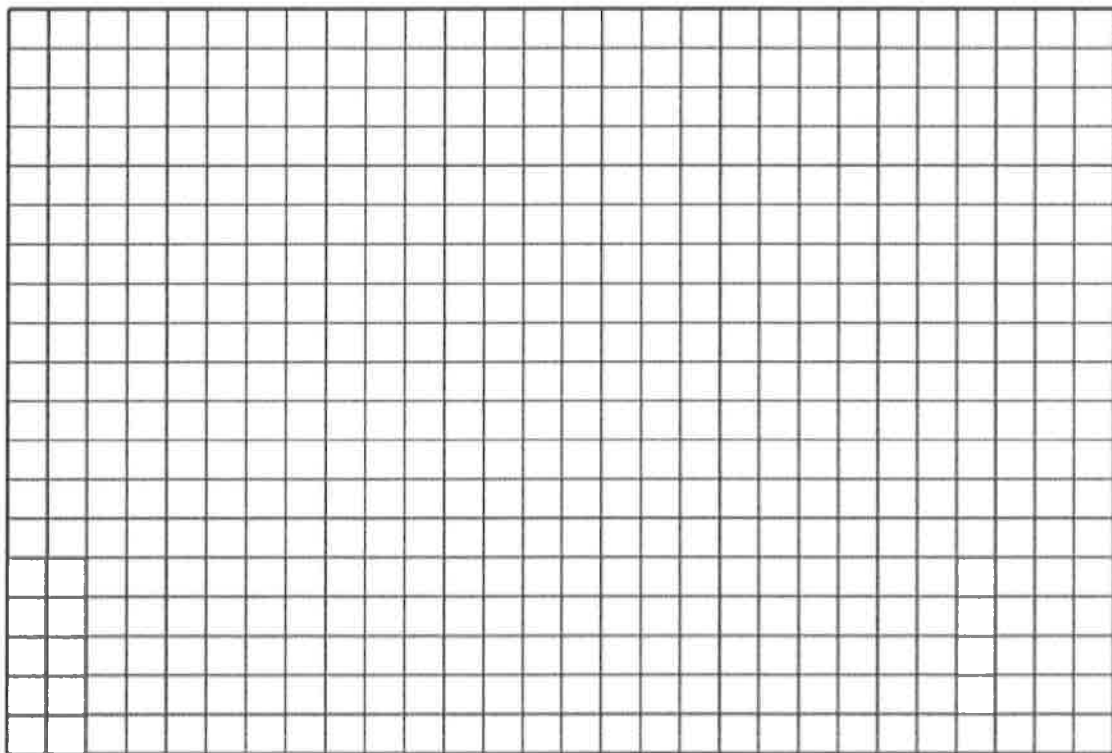
Sample Calculation:

Temperature, °C



Time in min.

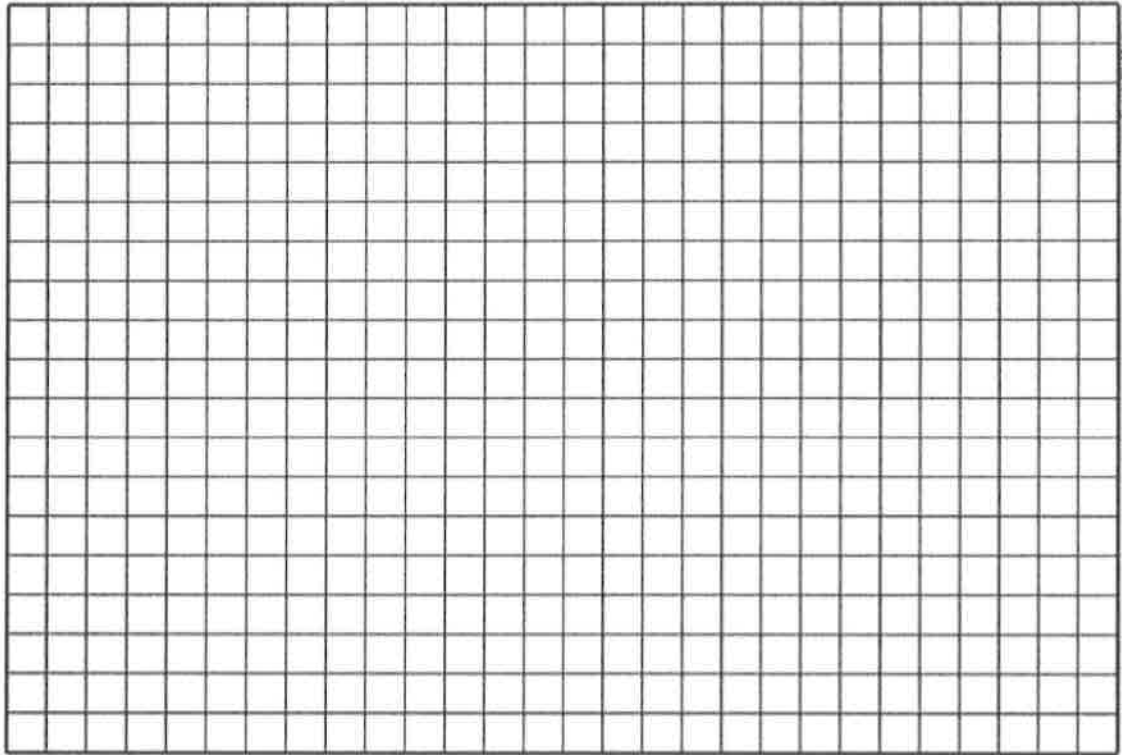
Temperature, °C



Time in min.

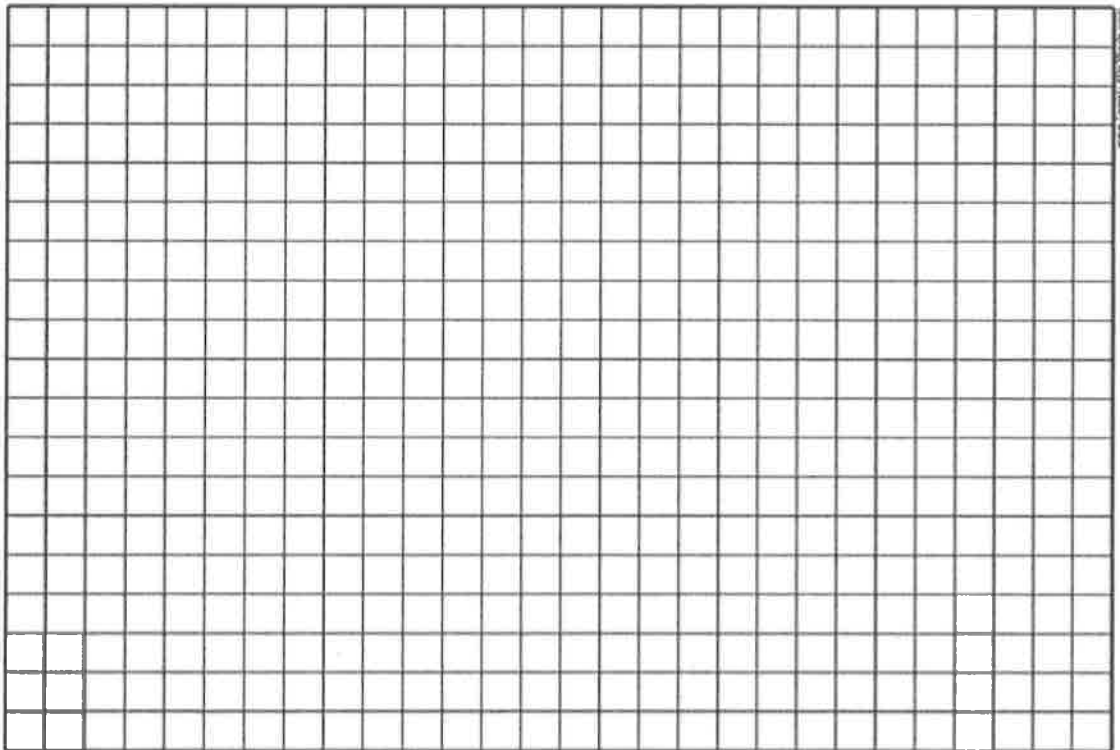


Temperature, °C



Time in min.

Temperature, °C



Time in min.