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CHEMISTRY WORKSHEET

CALORIMETRY

Course: Grade 12 Chemistry

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50 Practice Problems with Answers and Detailed Solutions

Full version of this worksheet contains 50 representative problems on:

✓heat absorbed or released using q = m·c·ΔT, ✓calorimeter heat capacity, ✓reaction vs surroundings sign convention, ✓molar enthalpy of reaction, ✓bomb calorimetry and combustion, ✓heat exchange and final temperature, ✓limiting reagent and unknown concentration.

All problems are designed to reflect style and difficulty commonly seen on quizzes, tests, and final exams. Detailed solutions are included to help students understand the reasoning behind every step.

PART A - FUNDAMENTAL PROBLEMS
Problems 1-15. Simple calculations and basic applications

PART B - INTERMEDIATE PROBLEMS
Problems 16-30. Typical unit test questions involving multiple steps

PART C - ADVANCED PROBLEMS
Problems 31-40. More challenging calculations requiring deeper understanding

PART D - CHALLENGE PROBLEMS
Problems 41-50. Difficult exam-style questions combining multiple concepts

SUITABLE FOR:
quizzes, unit tests and final exam preparation

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Section 1 - Essential Information – included in full version

This page includes only the formulas and habits needed to solve calorimetry problems efficiently.

Formula	Use
$q = m \times \Delta T$	Heat gained or lost by a substance
$\Delta T = T_f - T_i$	Temperature change with sign
$q_{cal} = C_{cal} \Delta T$	Heat absorbed by a calorimeter
$q_{reaction} = -q_{surroundings}$	Sign convention for coffee-cup and bomb calorimetry
$q_{hot} + q_{cold} = 0$	Heat exchange between substances
$\Delta H = q_{reaction}/n$	Molar enthalpy at constant pressure

Important definitions coffee-cup calorimetry is usually constant pressure, so q_p can be used for ΔH. Bomb calorimetry is constant volume, so the result is commonly treated as ΔU for SCH4U-style calculations.

Fast method identify what gains heat, identify what loses heat, convert volumes to masses when density is given, calculate ΔT carefully, include C_{cal} when provided, then apply the sign convention at the end.

Important notes for dilute aqueous solutions, use density = 1.00 g/mL and c = 4.184 J/(g·°C) only when stated or clearly assumed. For reactions, the solution and calorimeter are surroundings, while the chemical reaction is the system.

Common mistakes using only one solution volume after mixing, forgetting that $q_{reaction}$ has the opposite sign of q_{water}, ignoring the limiting reagent, mixing J and kJ, and using the sample mass instead of moles when a molar answer is requested.

Section 2 - Practice Problems

Part A - Fundamental Problems (Problems 1-15 - in full version)

Problem 8 When 0.725 g of ethanol, C_2H_5OH, burns in a bomb calorimeter with C_{cal} = 10.4 kJ/°C, the temperature increases by 2.19 °C. Calculate the molar enthalpy change for ethanol combustion

Part B - Intermediate Problems (Problems 16-30- in full version)

Problem 16 50.0 mL of 1.00 mol/L HCl is mixed with 50.0 mL of 1.00 mol/L NaOH in a coffee-cup calorimeter. The temperature rises from 22.0 °C to 28.8 °C. Assuming density = 1.00 g/mL and c = 4.184 J/(g·°C), calculate $\Delta H_{neutralization}$ in kJ/mol of H_2O formed

Part C - Advanced Problems (Problems 31-40 - in full version)

Problem 32 A 65.0 g metal sample at 95.0 °C is placed into 100.0 g of water at 20.0 °C. The final temperature is 24.6 °C. If only 92.0% of the heat lost by the metal is absorbed by the water, calculate the specific heat capacity of the metal

Part D - Challenge Problems (Problems 41-50 - in full version)

Problem 45 A coffee-cup calorimeter is calibrated by mixing 75.0 g of water at 60.0 °C with 75.0 g of water at 20.0 °C. The calorimeter starts at 20.0 °C, and the final temperature is 38.2 °C. A reaction is then carried out in the same calorimeter using 100.0 g of solution, and the temperature rises from 22.0 °C to 26.8 °C. Calculate C_{cal} and $q_{reaction}$ for the reaction trial

Section 3 - Solutions

Problem 8

Given:

m = 0.725 g ethanol, M = 46.07 g/mol, C_{cal} = 10.4 kJ/°C, ΔT = 2.19 °C.

Find: We need to determine molar enthalpy change in kJ/mol. This is the unknown quantity that must be calculated from heat transfer, calorimeter heat capacity, or the reaction energy balance

Method: Find the heat released by the sample, then divide by moles of ethanol

The calorimeter gives the heat released by the particular sample burned, but the question asks for a molar value. After finding $q_{reaction}$, divide by the moles of ethanol so the answer is in kJ/mol

Formula: $q_{cal} = C_{cal} \cdot \Delta T$, $q_{reaction} = -q_{cal}$, and $\Delta H°(molar) = q_{reaction}/n$

q_{cal} = (10.4 kJ/°C)·(2.19 °C) = 22.776 kJ;

n = 0.725 g / 46.07 g/mol

n = 0.0157 mol;

$\Delta H°(molar) = -22.776 \text{ kJ} / 0.0157 \text{ mol} = -1.45 \times 10^3 \text{ kJ/mol}$

The large negative molar value is expected for combustion of an organic fuel

Final answer: -1.45×10^3 kJ/mol

Problem 16

Given: V_{HCl} = 50.0 mL, c_{HCl} = 1.00 mol/L, V_{NaOH} = 50.0 mL, c_{NaOH} = 1.00 mol/L, m = 100.0 g, T_1 = 22.0 °C, T_2 = 28.8 °C.

Find: We need to determine $\Delta H_{neutralization}$. This is the unknown quantity that must be calculated from heat transfer, calorimeter heat capacity, or the reaction energy balance

Method: Calculate heat gained by the solution, then divide the reaction heat by moles of water formed

Strong acid and strong base neutralization forms water, so the molar enthalpy must be reported per mole of H_2O formed. The temperature rise gives heat absorbed by the solution, while the reaction heat has the opposite sign

Formula: $q_{solution} = m \cdot c \cdot \Delta T$, $q_{reaction} = -q_{solution}$, and $\Delta H = q_{reaction}/n$

ΔT = 6.8 °C; $q_{solution}$ = (100.0 g)·(4.184 J/(g·°C))·(6.8 °C);

n = (0.0500 L)·(1.00 mol/L)

$q_{solution}$ = 2845.12 J = 2.845 kJ;

n = 0.0500 mol H_2O;

ΔH = -2.845 kJ / 0.0500 mol = -56.9 kJ/mol

The value is close to the expected strong acid-strong base neutralization enthalpy

Final answer: -56.9 kJ/mol H_2O

Problem 32

Given: m_{metal} = 65.0 g, $T_{metal,i}$ = 95.0 °C, m_{water} = 100.0 g, $T_{water,i}$ = 20.0 °C, T_f = 24.6 °C, heat transfer efficiency = 92.0%.

Find: We need to determine c_{metal}. This is the unknown quantity that must be calculated from heat transfer, calorimeter heat capacity, or the reaction energy balance

Method: The water receives only 92.0% of the heat released by the metal, so divide the water heat by 0.920 before solving for c

The key detail is that the water receives only 92.0% of the heat lost by the metal. Therefore, the actual heat lost by the metal is larger than q_{water}, so the efficiency factor must be included in the equation

Formula: $q_{water} = 0.920[m_{metal} \cdot c_{metal} \cdot (T_{metal,i} - T_f)]$

$c_{metal} = [(100.0) \cdot (4.184) \cdot (24.6 - 20.0)] / [0.920(65.0) \cdot (95.0 - 24.6)]$

$c_{metal} = 0.457$ J/(g·°C)

Accounting for 92.0% efficiency increases the calculated c compared with assuming perfect transfer

Final answer: 0.457 J/(g·°C)

Problem 45

Given: calibration: 75.0 g hot water at 60.0 °C, 75.0 g cold water at 20.0 °C, T_f = 38.2 °C; reaction: m = 100.0 g, ΔT = 4.8 °C.

Find: We need to determine C_{cal} and $q_{reaction}$. This is the unknown quantity that must be calculated from heat transfer, calorimeter heat capacity, or the reaction energy balance

Method: First use hot and cold water mixing to find the calorimeter constant, then include that constant in the reaction heat calculation

The first water-mixing trial is a calibration experiment. The hot water loses heat, while the cold water and the calorimeter gain heat; once C_{cal} is known, the reaction trial includes both solution and calorimeter heat

Formula: $q_{hot\ lost} = q_{cold\ gained} + q_{cal}$; $q_{reaction} = -[m \cdot c \cdot \Delta T + C_{cal} \cdot \Delta T]$

$C_{cal} = \{[(75.0) \cdot (4.184) \cdot (60.0 - 38.2)] - [(75.0) \cdot (4.184) \cdot (38.2 - 20.0)]\} / (38.2 - 20.0)$

$C_{cal} = 62.1$ J/°C;

$q_{surroundings} = (100.0) \cdot (4.184) \cdot (4.8) + (62.1) \cdot (4.8) = 2306$ J = 2.31 kJ,

so $q_{reaction} = -2.31$ kJ

The reaction heat is negative because the solution and calorimeter warmed during the reaction trial

Final answer: $C_{cal} = 62.1$ J/°C; $q_{reaction} = -2.31$ kJ

Section 4 - Answer Key

Problem	Answer
8	-1.45×10^3 kJ/mol
16	-56.9 kJ/mol H_2O
32	0.457 J/(g·°C)
45	C_{cal} = 62.1 J/°C; $q_{reaction}$ = -2.31 kJ

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