Pre-AP Chemistry – Unit 10 – Thermochemistry

Objective 1: Students will understand and apply the 1st law of thermodynamics in the context of a chemical system. [KE.23.C.5]

Objective 2: Students will define and apply the concept of enthalpy to calculate heat exchange using constant pressure calorimetry. [KE.23.C.1, KE.23.C.4]

Objective 1

1. Define the following: system, surroundings, potential energy, kinetic energy, heat, exothermic, endothermic, 1st law of thermodynamics.
2. Where is energy in a chemical system stored?
3. Why is energy released or gained in a chemical reaction?
4. Which way does heat flow when exchanged?
5. List three common forms of energy.

Objective 2

1. Define the following terms: calorie, joule, heat capacity, specific heat, calorimetry, enthalpy, enthalpy of fusion, enthalpy of vaporization.
2. Calculate the heat gained when a 10g block of ice at –50 °C is converted to water at 82 °C.
3. A certain calorimeter is constructed and used to measure the specific heat of a metal. The 110g of water in the calorimeter was initially at 20 °C. If placing 250g of the metal, initially at 140°C, in the water caused its temperature to rise to 31 °C, then what is the specific heat of the metal?
4. The specific heat of copper is 0.385 J/g°C. How many joules of heat are necessary to raise the temperature of a 1.42kg block of copper from 25.0°C to 88.5°C?
5. How much energy is released as 15g of steam at 500K is converted to water at 350K?

Specific Heat Values	**Molar Heat of Fusion**
$H_2O_{(s)}$ = 2.09J/g°C	**H_2O = 6.01 kJ/mol**
$H_2O_{(l)}$ = 4.18 J/g°C	**Molar Heat of Vaporization**
$H_2O_{(g)}$ = 1.84 J/g°C	**H_2O = 40.6 kJ/mol**

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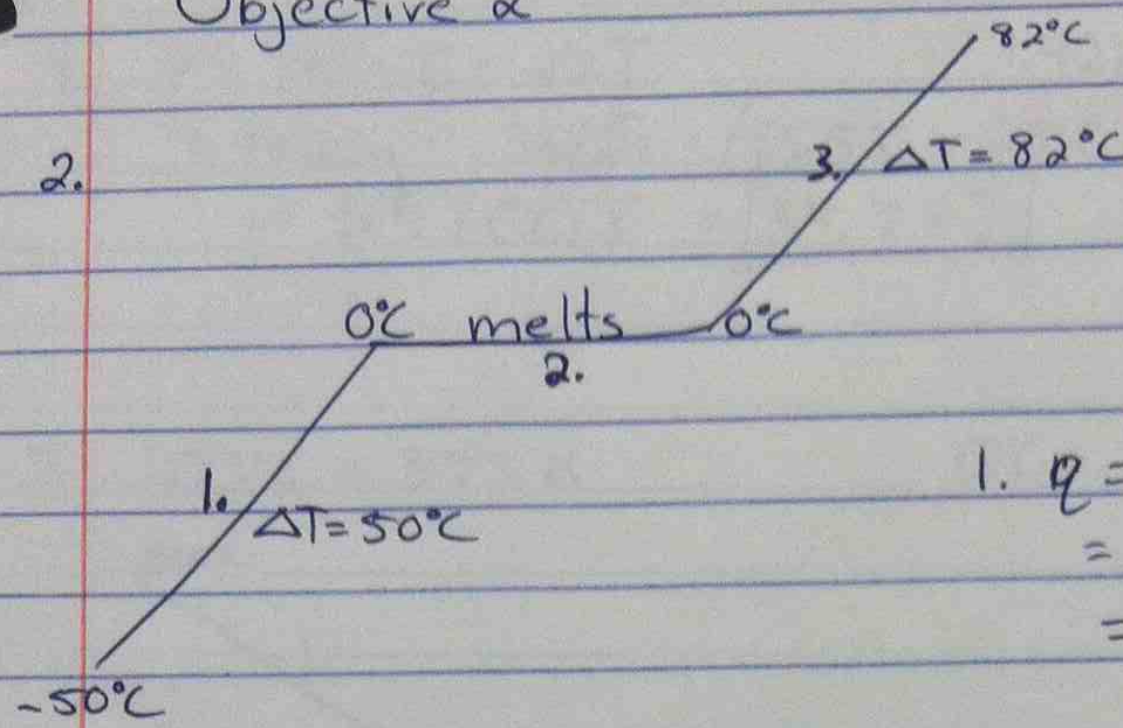
Objective 2



$10g \times \frac{1mol}{18g} = .556mol$

1. $q = m \cdot c \cdot \Delta T$
$= 10g \cdot 2.06 \frac{J}{g \cdot °C} \cdot 50°C$
$= 1030J = 1.03kJ$

2. $q = \Delta H_{fus} \cdot mol$
$6.01 \frac{kJ}{mol} \cdot .556 mol = 3.34kJ$

3. $q = m \cdot c \cdot \Delta T$
$= 10g \cdot 4.18 \frac{J}{g \cdot °C} \cdot 82°C$
$= 3427J = 3.43KJ$

$1.03kJ + 3.34kJ + 3.43kJ = \boxed{7.8kJ}$

3. $-q_{metal} = q_{water}$
$-m \cdot c \cdot \Delta T = m \cdot c \cdot \Delta T$
$-250g \cdot c \cdot (31°C - 140°C) = 110g \cdot 4.18 \frac{J}{g \cdot °C} \cdot (31°C - 20°C)$

$c = \frac{5057.8J}{250g \cdot 109°C} = \boxed{.186 \frac{J}{g \cdot °C}}$

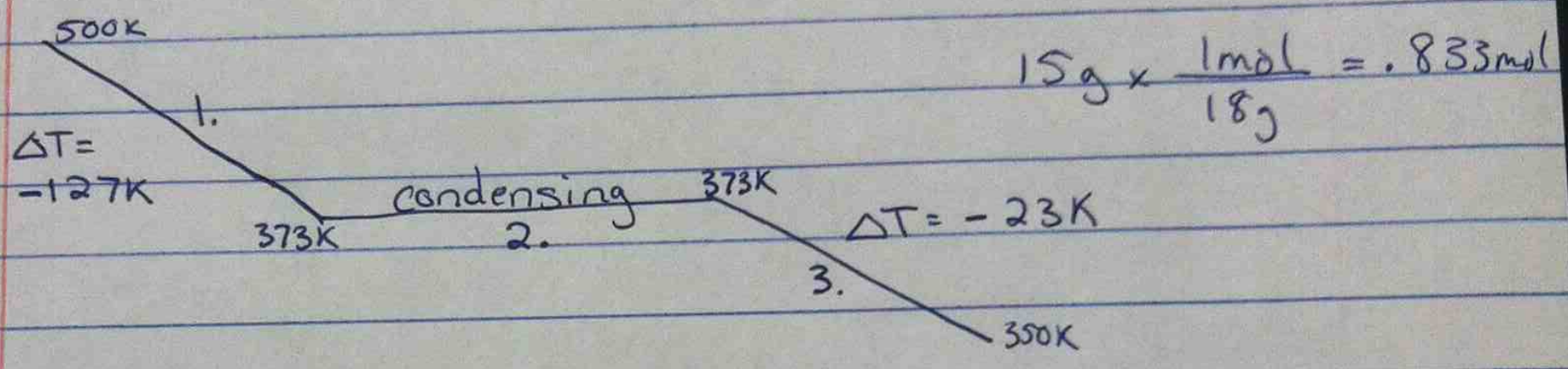
U10A HW

4. $q = m \cdot c \cdot \Delta T$ $\quad 1.42kg \times \frac{1000g}{1kg} = 1420g$

$= 1420g \cdot .385\frac{J}{g \cdot °C} \cdot (88.5°C - 25.0°C)$

$= 34700J = \boxed{34.7kJ}$

5. $100°C = 373K$ $\quad 0°C = 273K$



$15g \times \frac{1mol}{18g} = .833mol$

1. $q = m \cdot c \cdot \Delta T$

$= 15g \cdot 1.84\frac{J}{g \cdot K} \cdot -127K$

$= -3505J = -3.51kJ$

2. $q = \Delta H \cdot mol$ $\quad \Delta H$ is $-$ because it is condensing.

$= -40.6\frac{kJ}{mol} \cdot .833 mol$

$= -33.8kJ$

3. $q = m \cdot c \cdot \Delta T$

$= 15g \cdot 4.18\frac{J}{g \cdot K} \cdot -23K$

$= -1442J = -1.44kJ$

$-3.51kJ + -33.8kJ + -1.44kJ = \boxed{-38.8\ kJ}$