Name ____ Per. _________

Chem RG Chapter Packet 10 Causes of Change Mr. Nogales Score
Peer Review/corrected score

Assign	Section #	Name		
1.		Assignment Sheet printed (10 pts)		
2.		Notes 10.1 completed (10 pts)		
3.		Notes 10.2 completed (10 pts)		
4.		Worksheet 10.1 completed (10 pts)		
5.		Worksheet 10.2 completed (10 pts)		
6.		Ch 10 Test Review (10 pts)		
7.		Section and End of Chapter Summaries 20 pts (Total Points = 8 x 10 = 80)		

Notes:

1. Your lab report is turned in by itself and receives a separate grade.
2. Section and End of Chapter summaries are part of your grade.
3. Please note, that if a problem requires you to show work and you do not, you will not receive credit. therefore, SHOW YOUR WORK!

EVIDENCE (after you take notes.) You should have at least 4 types of evidence for each set of notes.

1. Number new concepts	1,2,3…/A,B,C…	2. Delete/Cross out unimportant information	~~Unimportant~~
3. Circle vocab/key terms	(Key Terms)	4. Identify points of confusion	?
5. Underline/Highlight main Ideas	Main Ideas	6. Identify information to be used on a test, essay…	*
7. Fill in gaps of information. Reword or paraphrase.	^	8. Create visuals/symbols of important information	Visuals/symbols

Experiment A16 Temperature of a Bunsen Burner Flame.

Objectives

- **Observe** the change in temperature of a known mass of water when a heated metal object of known mass is placed in it.
- **Use** the specific heat of the metal and of water to calculate the initial tempera- ture of the object.
- **Relate** the temperature of the object to the temperature of the Bunsen burner flame.

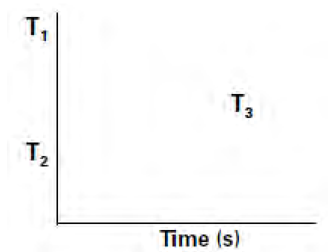
Introduction

When a hot solid is immersed in a cool liquid, heat flows from the hot object to the cool liquid. In fact, the number of joules of energy lost by the hot solid (ΔQ_1) equals the number of joules of energy gained by the cool liquid (ΔQ_2).

$$-\Delta Q_1 = +\Delta Q_2$$

Heat lost or gained depends on the object's mass and specific heat. Every material has a specific heat = number of joules of energy needed to change the temperature of one gram of material one Celsius degree. If the mass, temperature change, and specific heat of a substance are known, the heat lost or gained can be calculated using this relationship: $\Delta Q =$ (mass)(specific heat)(Δt). The temperature change is defined as $\Delta t = t_{final} - t_{initial}$.

In this experiment, you determine the temperature of a Bunsen burner flame by heating a sample of a known metal and then immersing the hot metal at temperature t_1 into a measured quantity of water at temperature t_2 . As heat flows from the hot metal to the cool water, the two materials approach an intermediate temperature t_3 . These changes are shown graphically.



$$-\Delta Q_1 = \Delta Q_2$$

$$-(m_1)(c_{p1})(t_3 - t_1). = (m_2)(c_{p2})(t_3 - t_2)$$

By solving the equation for t_1 , the initial temperature of the hot metal can be found. The metal should be the same temperature as the Bunsen burner flame.

Materials

balance, Bunsen burner, foam cup, iron ring, 20-30 g metal (Fe, Zn or Cu), Nichrome wire, ring stand, thermometer

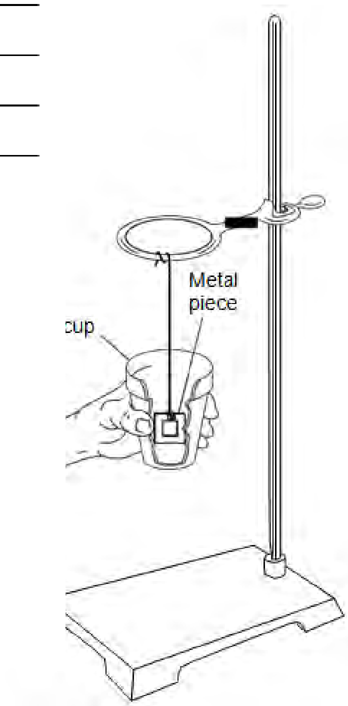
Procedure: Caution – the metal will get very hot.

1. Obtain the mass of an empty foam cup to the nearest 0.01 g, and record the mass in the Data Table.
2. Fill the cup about two-thirds full with water. Obtain the mass of the cup and water and record it in the Data Table.
3. Read and record the temperature of the water to the nearest 0.1°C.
4. Obtain the mass of the metal object to the nearest 0.01 g and record it in the data table.
5. Use the Nichrome wire to attach the metal object to the iron ring clamped on the ring stand. The object should hang about 7 or 8 cm below the ring.
6. Adjust the Bunsen burner so that it produces a medium hot, blue flame. Move the flame under the metal object, and heat it for about 5 minutes.
7. Turn off Bunsen burner and move it aside. SLOWLY lift the foam cup with the water so that the hot metal becomes immersed in the water, as shown in Figure B. Hold the cup in this position for 1 min and then read and record the temperature of the water in the Data Table.

Data Table	
Type of metal used	
Mass of metal	g
Specific heat of metal (from Table 1)	J/g °C
Mass of empty foam cup	g
Mass of foam cup and water	g
Initial temperature of water	°C
Final temperature of water	°C

Table 1: Specific Heats of Materials

Substance	Specific heat (J/g °C)
Copper or Zinc	0.385
Iron	0.444
Manganese	0.481
Nickel	0.470
Platinum	0.131
Tungsten	0.134
Water	4.18



Calculations

1. Calculate the change in the water temperature.
2. What is the mass of the water that was heated by the metal object?
3. Calculate the heat gained by the water. (Hint: Use $q = m \cdot \Delta t \cdot Cp$ and solve for q)
4. Calculate the temperature of the hot metal object. (Hint: Refer to the Introduction. The heat lost by the metal equals the heat gained by the water.)
    - Use this formula: $-(m_1)(C_{p1})(t_3 - t_1) = (m_2)(C_{p2})(t_3 - t_2)$ where t_1 is your unknown metal temperature, t_2 is the original water temperature and t_3 is the same <u>final</u> temperature for water and metal.
    - Put your metal data on the left side (don't forget the negative sign) and your water data on the right side. Then solve for t_1. This is the metal temperature when you heated it in the flame.
5. What is the temperature of the Bunsen burner flame?

Question

1. a. Using the specific heats in Table 1 above, which metal would raise the water temperature the most?

   b. Which metal would raise the water temperature the least?

General Conclusion

1. In the eighteenth century, clothes were pressed using the heat from a heavy piece of metal with a flat, smooth side. This solid metal "iron" had to be heated periodically on top of a wood- or coal-burning stove. Disregarding cost, which of the metals listed in Table 1 would be the best to use for this purpose? Explain your answer.

Chapter 10 - Causes of Change

10.1 Energy Transfer

- Define enthalpy
- Distinguish between heat and temperature
- Perform calculations using molar heat capacity.

Temperature is different than heat

Temperature measures the ________ ________ ________of the particles in a sample of matter.

Heat - the energy transferred from ________things to ________things.

A drop of boiling water hurts, but a Kg of boiling water kills (________temperature, but more ________)

Enthalpy

Enthalpy - total energy content of something.

________ measurement is calories (food). Now we use ________
Example - When 1 gram of water absorbs 4.184 joules of energy (heat) its temperature goes up 1 ºC

Units of heat - calories or Joules

1 calorie = amount of heat needed to raise the temperature of ________ gram of water by ________

It takes 4.184 J to do the same thing. So, 1 calorie = ________ A food Calorie is really a kilocalorie

Energy conversions

How much energy is needed to heat 15 g of water by 25ºC? We'll do in J & cal.

15 g • 25ºC • 4.184 J/gºC = 1569 ________ or 15 g • 25ºC • 1 cal/gºC = 375 ________

Convert 10 cal to J . . .
10 cal • 4.184 J/cal = 41.84 J

Convert 10 J to cal . . .
10 J • 1 cal/4.184 J = 2.39 cal

Use above to help with online HW.

Some Equalities for Heat

Heat is measured in calories or joules
1 kcal = 1000 cal = 1 Cal (a food calorie). 1 calorie = 4.184 J. 1 kJ = 1000 J

Energy and Nutrition

1 Calorie (nutritional) = 1 kcal. 1 Cal = 1000 cal
Caloric Food Values: Carbohydrate = 4 kcal/g. Fat = 9 kcal/g. Protein = 4 kcal/g
1.0 cup of whole milk contains 12 g of carbohydrate, 9.0 g of fat, & 9.0 g of protein. How many kcal (Cal)?

Solution:	12 g carbo x 4 kcal/g	=	48 kcal
	9.0 g fat x 9 kcal/g	=	81 kcal
	9.0 g protein x 4 kcal/g	=	36 kcal
	Total kcal	=	165 kcal

Specific Heat

Different substances have different ________ for storing energy

It may take 20 minutes to heat water to 75°C. However, the same mass of aluminum might require ________ minutes and the same amount of ________ may take only 2 minutes to reach the same temperature.

Learning Check

When you heat 200 g of water for 1 minute, the water temperature rises from 10°C to 18°C. If you heat 400 g of water at 10°C in the same pan with the same amount of heat for 1 minute, what would you expect the final temperature to be? 1) 10 °C 2) 14°C 3) 18°C. Solution = 2) 14°C

Heating ______ the mass of water using the ______ amount of heat raises temp only ________ as much.

Specific Heat Values

Specific heat is the heat needed to raise the temperature of ________ g of a substance by ________ °C

	cal/g°C	J/g°C
water	1.00	4.184
aluminum	0.22	0.90
copper	0.093	0.39
silver	0.057	0.24
gold	0.031	0.13

Learning Check

A substance with a large specific heat ___________________________

When ocean water cools, the surrounding air ___________________________

Sand in the desert is hot in the day, and cool at night. Sand must have a ___________________________

Some things heat up easily. Some take a great deal of energy to change their temperature.
Specific Heat Capacity - amount of heat to change the temperature of 1 g of a substance by 1ºC

Molar & Specific Heat

Table of molar heat on p. 343. Also table of specific heat above.
Water has a high specific heat: 75.3 J/K•mol or 4.184 J/gºC

Check your ________!
The amount of heat it takes to heat something is the same as the amount of heat it gives off when it cools because...Law of conservation of ________.

Molar Heat Capacity

The amount of heat necessary to raise the temperature of 1 ________ of the substance 1 ________
Every substance has its own special value.
Abbreviated as ________. ________ Heat of Water = 75.3 J/mol•K

Specific Heat Capacity

The amount of heat necessary to raise the temperature of 1 gram of the substance 1________
Specific Heat of Water = Cp = 4.184 J/gºC

Molar Heat
$q = n \cdot \Delta T \cdot Cp$
where q => heat, J
n => moles
ΔT = change in temperature, K
Cp => specific heat, J/mol•K

Specific Heat
$q = m \cdot \Delta t \cdot Cp$
where q => heat, J
m => mass, g
Δt = change in temperature, ºC
Cp => specific heat, J/g•ºC

Be sure to _______if the problem is in grams or moles! The Cp values will be different.

________ Heat Calculations

A hot-water bottle contains 41.6 moles of water at 338 ________. If the water cools to body temperature (310 K), how many joules of heat could be transferred to sore muscles? $q = n \cdot \Delta T \cdot Cp$ (1st calculate ΔT)

Heat = q = m x ΔT x Cp (H_2O)

41.6 mol x 28 K x 75.3 J/mol•K = 88 000 J (remember significant figures)

= 88 kJ (since 1 kJ/1000 J)

________ Heat Calculations

A hot-water bottle contains 750 g of water at 65°C. If the water cools to body temperature (37°C), how many joules of heat could be transferred to sore muscles? $q = m \cdot \Delta t \cdot Cp$

Heat = q = m x ΔT x Cp (H_2O)

750 g x 28°C x 4.18 J = 88 000 J (remember significant figures)

= 88 kJ (since 1 kJ/1000 J)

Now You Try It – SHOW YOUR CALCULTIONS BELOW

It takes 1950 joules to heat 10.0 mol of a metal from 295 K to 302 K.
What is its molar heat capacity? What metal is it? Hints: use $q = m \cdot \Delta T \cdot C_p$, solve for C_p, then use table on p. 343.
Cp =

Metal =

Now You Try It – SHOW YOUR CALCULATIONS

Iron has a Cp of 0.449 J/g°C. How much heat to change the temperature of 48.3 g of iron by 32.4°C?

Heat Transfer

$q_{lost} = - q_{gained}$

$(m \cdot \Delta t \cdot Cp)_{lost} = - (m \cdot \Delta t \cdot Cp)_{gained}$

You will use this principle in lab A16. Just follow the example, don't take notes

If 100. g of iron at 100.0°C is placed in 200. g of water at 20.0°C in an insulated container, what will be the final temperature, °C, of the iron & water when both are at the same temperature? Iron specific heat is 0.106 cal/g°C.

$(100.g \cdot 0.106 cal/g^oC \cdot (T_f - 100.)^oC) = q_{lost}$ $- q_{gained} = (200.g \cdot 1.00 cal/g^oC \cdot (T_f - 20.0)^oC)$

$[10.6(T_f - 100.^oC)]_{lost}$ $= [- 200.(T_f - 20.0^oC)]_{gained}$

$[10.6T_f - 1060^oC]_{lost}$ $= [- 200.T_f + 4000^oC]_{gained}$

Collect like terms . . .

$(10.6 + 200.)T_f = (1060 + 4000)^oC$

$T_f = (5060/211.)^oC = 24.0^oC$

10.2 Using Enthalpy

- Define thermodynamics
- Understand thermodynamic equations as being endothermic or exothermic

Energy

Energy is measured in ________. Every reaction has an energy ________ associated with it

Exothermic reactions ________ energy, usually as heat. Endothermic reactions ________ energy

Energy is <u>stored</u> in ________ between atoms

An equation that includes energy is called a ________ equation. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 802.2$ kJ
1 mole of CH_4 <u>makes</u> 802.2 kJ of energy. When make 802.2 kJ of energy, also make 2 moles of water

Exothermic *vs.* Endothermic

This is __________ because it makes energy. $CH_4 + 2O_2 \rightarrow CO_2 + 2\ H_2O + 802.2$ kJ $\Delta H = -802.2$ kJ

The opposite reaction is ____________. $CO_2 + 2H_2O + 802.2$ kJ $\rightarrow CH_4 + 2O_2$ $\Delta H = +802.2$ kJ

Enthalpy

The heat content a substance has at a given temperature and pressure
Can't be measured directly because there is no set starting point
The reactants start with a heat content. The products end up with a heat content

Symbol is H. Change in enthalpy is ΔH. Called, "________ H"

If heat is ________ the heat content of the <u>products</u> is <u>lower</u>. ΔH is negative (________)

If heat is <u>absorbed</u> the heat content of the <u>products</u> is ________. ΔH is positive (endothermic)

Exothermic
Products are ________ in energy than the reactants.
Releases energy.
ΔH is (-).

Endothermic
Products are________ in energy than the reactants
Absorbs energy
ΔH is (+)

Heat of Reaction

The heat that is released or absorbed in a chemical reaction. Equivalent to ΔH
$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} + 393.5$ kJ. Also shown as $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H = -393.5$ kJ
In a thermochemical equation it is important to say what the states are.
$H_2(g) + 1/2O_2\,(g) \rightarrow H_2O(g)$ $\Delta H = -241.8$ kJ for gaseous water
$H_2(g) + 1/2O_2\,(g) \rightarrow H_2O(l)$ $\Delta H = -285.8$ kJ for ________ water

Calorimeter - measure the energy content of something by change in temperature of the water (<u>test question</u>)

Entropy (randomness)

A solid has an ________ arrangement.

A liquid has the molecules moving ________ to each other.

A gas has molecules moving ________ ________ the place.

Chem RG Chapter 10 worksheet. SHOW YOUR WORK!

10.1 Book Reference: See sample problem "A" p. 342 to learn how to do the problems below, which are p. 342 #1-4 and p. 344 #8-13.

q = nCΔT where q = heat (in J), C = molar heat capacity (in J/mol·K) and *Δ?T* = change in temperature (^{o}C or K). Use this equation for these problems.
(Now, if instead of mol you are given g, then use *q = mΔTCp* where m = mass and Cp = specific heat capacity. You'll see this on your Moodle quizzes).

1. The molar heat capacity of tungsten is 24.2 J/K•mol. Calculate the energy as heat needed to increase the temperature of 0.40 mol of tungsten by 10.0 K.

2. Suppose a sample of NaCl increased in temperature by 2.5 K when the sample absorbed 1.7×10^2 J of energy as heat. Calculate the number of moles of NaCl if the molar heat capacity is 50.5 J/K•mol.

3. Calculate the energy as heat needed to increase the temperature of 0.80 mol of nitrogen, N_2, by 9.5 K. The molar heat capacity of nitrogen is 29.1 J/K•mol.

4. A 0.07 mol sample of octane, C_8H_{18}, absorbed 3.5×10^3 J of energy. Calculate the temperature increase of octane if the molar heat capacity of octane is 254.0 J/K•mol.

8. Calculate the molar heat capacity of diamond, given that 63 J were needed to heat a 1.2 g of diamond by 1.0×10^2 K.

9. Use the molar heat capacity for aluminum from **Table 1** to calculate the amount of energy needed to raise the temperature of 260.5 g of aluminum from 0°C to 125°C.

10. Use the molar heat capacity for iron from **Table 1** to calculate the amount of energy needed to raise the temperature of 260.5 g of iron from 0°C to 125°C.

11. A sample of aluminum chloride increased in temperature by 3.5 K when the sample absorbed 1.67×10^2 J of energy. Calculate the number of moles of aluminum chloride in this sample. Use **Table 1**.

12. Use **Table 1** to determine the final temperature when 2.5×10^3 J of energy as heat is transferred to 0.20 mol of helium at 298 K.

13. Predict the final temperature when 1.2 kJ of energy as heat is transferred from 1.0×10^2 mL of water at 298 K.

Hint for #13. Change kJ to joules. Change mL to g (use density of water = 1 g/ml). Then find delta T (the change in temperature). Then apply that change to the initial temperature (298 K) to get the final temperature.

Try to get these answers: p. 342: 1) 96.8 J. 2) 1.3 mol. 3) 221.16 J. 4) 196.85 K
Now, do p. 344: 8) 6.3 J/K•mol. 9) 29.2 kJ. 10) 14.6 kJ. 11) 0.52 mol. 12) 358 K. 13) 295 K.

10.2 worksheet. Use the formula $\Delta H = C \cdot \Delta T$ & look on page 343 to find the value of C.

See sample problems "B" & "C" p. 346

p. 346:1-3 SHOW YOUR WORK!

1. Calculate the molar enthalpy change of $H_2O(l)$ when liquid water is heated from 41.7°C to 76.2°C (ans. 2600 J/mol)

2. Calculate the ΔH of NaCl when it is heated from 0.0°C to 100.0°C. (ans. 5050 J/mol)

3. Calculate the molar enthalpy change when tungsten is heated by 15 K. (ans. 360 J/mol)

p. 347:1-3

1. The molar heat capacity of Al(s) is 24.2 J/K·mol. Calculate the molar enthalpy change when Al(s) is cooled from 128.5°C to 22.6°C. (ans. -2560 J/mol, (-) since cooling)

2. Lead has a molar heat capacity of 26.4 J/K·mol. What molar enthalpy change occurs when lead is cooled from 302°C to 275°C? (ans. -713 J/mol)

3. Calculate the molar enthalpy change when mercury is cooled 10 K. The molar hat capacity of mercury is 27.8 J/K·mol. (ans. -280 J/mol)

Ch 10: TEST Review 10 points Name: ___________________________________ Period: ________

What Symbol (q, H, m, n, Δ) stands for Change in: _____. Enthalpy: _____. Energy as Heat: _____. mol: _____. mass: ______.

Using a calorimeter is possible due to the known specific heat of ______________

Assign each of the following to the appropriate Exothermic or endothermic boxes below:

l→ s, l→ g, g → l, g → s, s → g, s→ l, condensation, deposition, melting, sublimation, evaporation, freezing, +ΔH , -ΔH,

feels warm, feels cold, particles get less orderly, particles get more orderly, releases energy, absorbs energy,

energy is reactant, energy is product, products at higher energy than reactants, products at lower energy than reactants

Exothermic:	Endothermic:

If you measure the average kinetic energy of a substance you are also measuring the substance's ________________________

If you measure the energy transferred to a sample you are measuring the _____________

To convert Kelvin to Celsius you (add/subtract) 273, and to convert Celsius to Kelvin you (add/subtract) 273.

Therefore, 234 K = _________ ˚C and 456 ˚C = _________ K, Which temperature is higher, 250 ˚C or 400 K?________

TWO POINTS EACH – SHOW CALCULATIONS. For below use q = nΔTCp if starting with moles & q = mΔTCp if starting with grams.
Cp: water = 4.18 J/gºC or 36.8 J/nK. Ice = 2.06 J/gºC for ice; Steam = 2.02 J/gºC. ΔH_{fus} = 334 J/g. ΔH_{vap} = 2060 J/g.

Calculate how much energy (heat) is needed to raise the temperature of 4.5 grams of Copper (C_p = 0.385 J/g • K) by 47 K.

Calculate how much energy (heat) is needed to raise the temperature of 4.5 moles of Helium (C_p = 20.8 J/K • mol) by 47 K.

Calculate the specific heat of a substance if it takes 6.7 J of energy to raise the temperature of 2.3 grams of the substance by 78 K.

Calculate the molar heat of a substance if it takes 6.7 J of energy to raise the temperature of 2.3 moles of the substance by 78 K.

SHOW YOUR WORK! SHOW YOUR CALCULTIONS ABOVE !