

# UWO **CHEM 1302**

**Winter 2025, Chapter 3 Notes**



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# 3. Heat, Work, and Energy

## 3.1 Introduction to Thermodynamics

3.1.1

### Key Terms

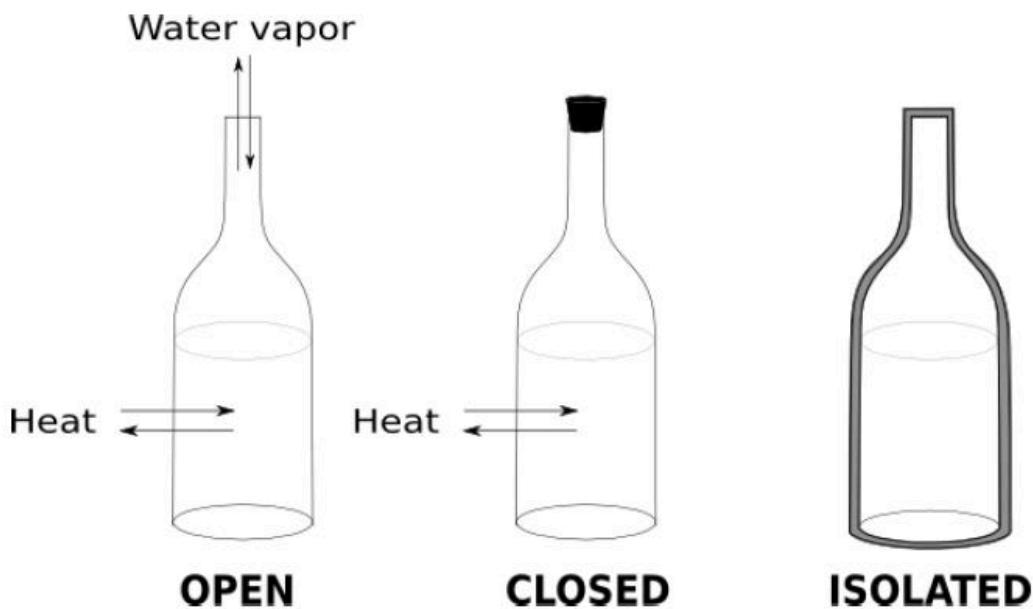
The following lesson is low yield in terms of exams. We'll just focus on **understanding**.💡

#### System Vs Surroundings

**System**-reactants and products that are part of the **chemical reaction**

Term	Definition
System	The collection of matter that is under consideration
Surroundings	Everything that is not a part of the system
Universe	System + Surroundings
Open system	Can exchange matter and energy with surroundings
Closed system	Can exchange only energy with surroundings
Isolated system	Cannot exchange matter or energy with surroundings (adiabatic, $q = 0$ )

## Open Vs Closed Vs Isolated Systems



### Examples:

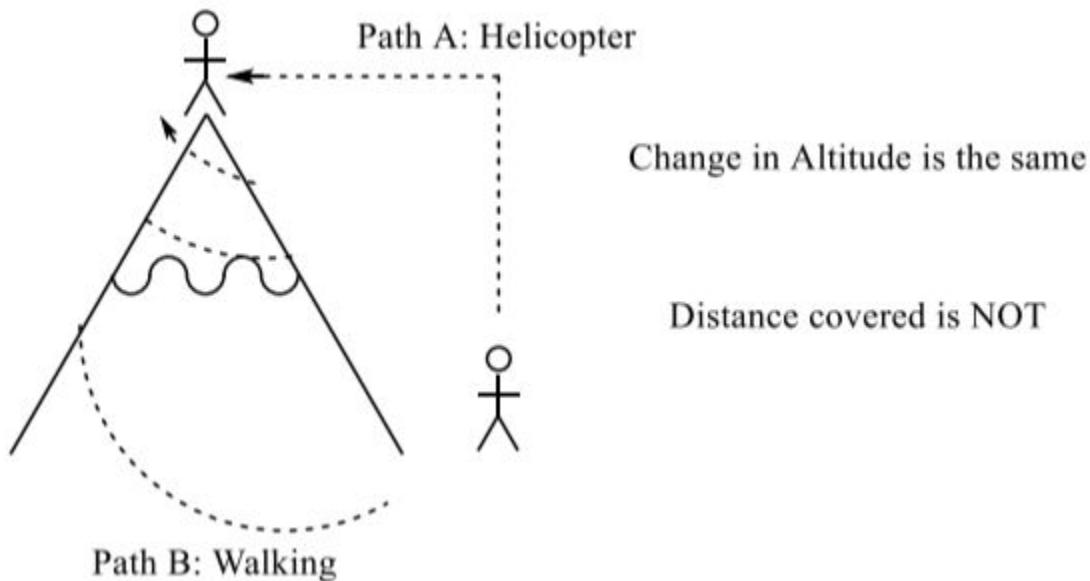
- open system → cup of coffee
- closed system → bottle of pop
- isolated system → perfect thermos

## State Variables

The following lesson is low yield in terms of exams. We'll just focus on **understanding**.💡

**State variable** (aka state function): is a variable that depends only on the current state of a system, NOT on the path taken to that state.

*Example:*



Is the distance travelled or change in altitude a state function?

the change in altitude is a state function because no matter which path was taken .

## Example: State Variables

Which of the following variables are state variables? Label each as yes or no.

a) mass yes

b) temperature yes

c) work No

d) pressure yes

e) volume yes

f) density yes

## Extensive vs. Intensive Properties

The following lesson is **low yield** in terms of exams. We'll just focus on **understanding**. 

**Extensive properties:** depend on the **extent (or amount)** of a substance

**Examples:** mass, volume, amount of heat released in a combustion reaction

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**Intensive properties:** does not depend on the amount of substance

*Examples:* melting point, boiling point, density

Think\* no matter how much water you are boiling in a pot, water will always boil at 100°C and melt at 0°C!

**When we take a ratio of 2 extensive properties what do we get?**

an intensive property

*Example:* mass/volume=density

## Additional Terminology

**Reversible Process:** The system is at equilibrium (and the reaction can go in either direction).

*Example:*  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$

**Irreversible Process:** The system is not at equilibrium and proceeds in one direction only. This reaction is spontaneous in the forward direction.

*Example:*  $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$  (combustion reaction)

**Spontaneous Process:** An irreversible process that has a predetermined direction.

*Example:* An egg shell breaks during birth

---

**Isothermal:** There is no change in temperature for the system

$$\Delta T = 0$$

For Ideal Gases:  $\Delta U = 0, \Delta H = 0$

**Isobaric:** There is no change in pressure

$$\Delta p = 0$$

**Isochoric:** There is no change in volume

$$\Delta V = 0$$

**Adiabatic:** No heat exchange (no heat enters or leaves the system)

$$q = 0$$

**Exothermic:** A Process in which heat is release (absorbed/released) to the surrounding (system/surroundings)

$$q < 0$$

**Endothermic:** A Process in which heat is absorbed (absorbed/released) into the the system (system/surroundings)

$$q > 0$$

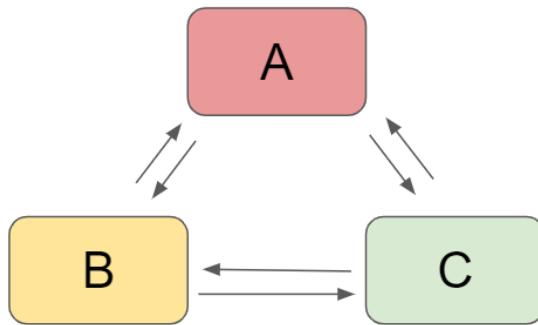
**Standard Conditions:** Pressure of all gases is 1 atm, all solutions have 1M concentration, and temperature is 298K (25C). This is denoted as  $H^\circ S^\circ G^\circ$  etc.

## 3.2

# 0th Law of Thermodynamics

3.2.1

## 0<sup>th</sup> Law of Thermodynamics



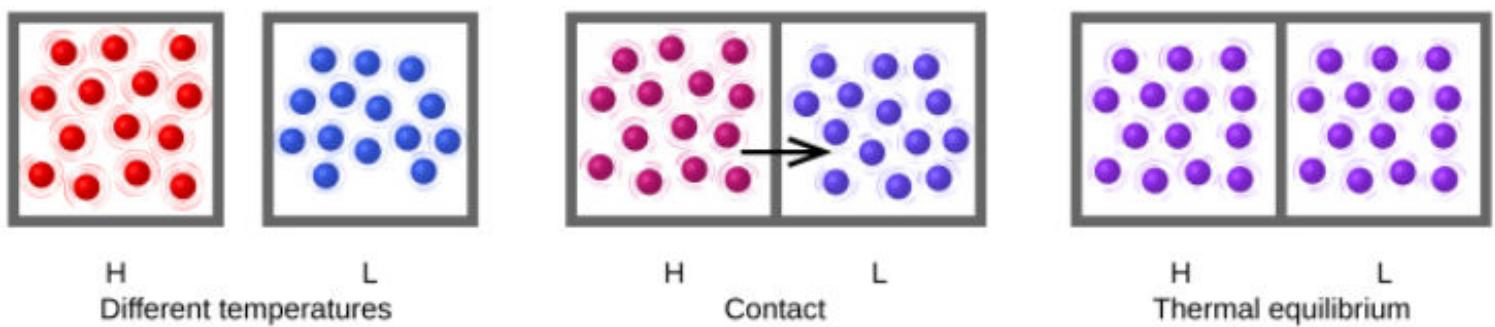
- **0th law:** If a body, A, is in thermal equilibrium with a second body, B, and that body is in thermal equilibrium with a third body, C, then A and C are in **thermal equilibrium**.
- If two bodies/systems are in **thermal equilibrium** with one another, it means they are at the **same temperature!**

## What is the difference between thermal equilibrium and temperature?

**Thermal equilibrium:** When two substances are in **physical contact** and **exchange no heat energy** with one another.

**Temperature:** A measure of the **average kinetic energy** of the particles in a sample of matter

Increasing the temperature means that average kinetic energy of the particles will (increase/decrease) increase.



### 3.3

## Relating Calorimetry to the 1st Law

3.3.1

### Calorimetry

- We use calorimetry to measure the heat flow of a system. The following equation can be used to calculate heat flow using given that  $C$  is the **heat capacity** of the calorimeter

$$q = C\Delta T$$

- Heat capacity,  $C$ , is dependent on the substance and is different depending on whether the calorimetry is performed at constant volume,  $C_V$  or constant pressure,  $C_P$
- Specific heat capacity is expressed as a function of mass (  $C_s$  or  $c$  )
- Molar heat capacity is expressed as a function of the number of moles (  $C_m$  ).

---

### Case 1: Constant Volume Process, $\Delta V = 0$

$$\Delta V = 0 \therefore w = 0 \therefore \Delta U = q_v$$

$$\Delta U = q_v = \int_{T_1}^{T_2} nC_{V,m} dT = nC_{V,m}(T_f - T_i)$$

Note that we can use the exact same equations if we are told there is free expansion into a vacuum:

- When expanding freely into a vacuum there is **no external pressure**
- Since  $w=-P\Delta V$  and  $P=0$ , then  $w=0$

$$p_{ex} = 0 \therefore w = 0$$

$$\Delta U = q + w \therefore \Delta U = q_v$$

---

## Case 2: Constant Pressure Calorimetry, $\Delta p = 0$

We need to define a new state function to keep track of heat exchange at constant  $p$

$$H = U + pV$$

$$\Delta H = q_p = \int_{T_1}^{T_2} nC_{p,m}dT = nC_{p,m}(T_f - T_i)$$

It can be shown that

$$C_{p,m} = C_{V,m} + R$$

### Case 3: Isothermal Conditions

Under isothermal conditions (constant temperature),  $\Delta T=0$

 WIZE TIP

$\Delta U$  in the equation:  $\Delta U = q + w$ , depends only on temperature!

So when there is no temperature change under isothermal conditions, there is no change in  $U$  ( $\Delta U=0$ )

**Note:** because it's isothermal,  $\Delta U=0$

$$\Delta U = q + w = 0 \therefore q = -w$$

## Heat Capacity of Ideal Gases

- It can be shown that

$$C_{p,m} = C_{V,m} + R$$

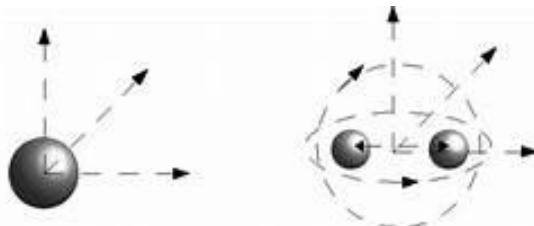
- For an ideal monoatomic gas

$$C_{V,m} = \frac{3}{2}R \quad C_{p,m} = \frac{5}{2}R$$

- For an ideal diatomic gas

$$C_{V,m} = \frac{5}{2}R \quad C_{p,m} = \frac{7}{2}R$$

The following is **low yield** info that can help you memorize these equations (but they will most likely be provided to you on a formula sheet).



Monatomic: only translational degrees of freedom.

Diatom: translational, rotational, and vibrational degrees of freedom.

- Monoatomic gases have 3 degrees of freedom (they are able to move in 3 directions)
- While diatomic gases have 5 degrees of freedom (they are able to move in 5 directions)
  - The degrees of freedom matches the numerator of the Cv equation for each!

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### 3.3.3

The enthalpy of combustion for ethanol  $C_2H_5OH$  is -1370.7 kJ/mol. Calculate q and  $\Delta H$  when 45 g of ethanol are burned at 1 atm and 298 K.

a.  $q$  (kJ)

b.  $\Delta H$  (kJ)

(Don't include units in your answer and round to the nearest whole number)

a.

-1337 kJ

---

b.

-1337 kJ

---

$$n_{C_2H_5OH} = 45g \times \frac{1\text{ mol}}{46.068\text{ g}} = 0.976\text{ mol}$$

$$q = n \Delta H = 0.976 \times -1370.7 = -1337\text{ kJ}$$

at constant pressure

$$q = \Delta H$$

## Practice: Isothermal Expansion

4 moles of neon was confined to a 8L flask initially at room temperature underwent an isothermal expansion into a vacuum at 348K. Calculate  $\Delta U$ ,  $\Delta H$ , and  $q$  for this process.

Neon (monatomic ideal gas)

$$n = 4 \text{ mols} \quad T = 384 \text{ K}$$

$\Delta U = 0$  for an isothermal process the internal energy is only depends on temperature

$\Delta H = 0$  enthalpy is also temperature-dependent on ideal gas

$W = 0$  in a vacuum no external pressure oppose the gas

$$\Delta H^{\circ} = q + W^{\circ}$$

$$q = 0$$

## 3.4 The 1st Law of Thermodynamics and Work

3.4.1

### The 1<sup>st</sup> Law of Thermodynamics

- **1<sup>st</sup> Law of Thermodynamics:** The energy of an isolated system is constant,  $\Delta E = 0$ .
  - Sometimes this is re-worded as "energy cannot be created or destroyed".
- $E$  is very difficult to measure so instead we measure changes in  $E$  indirectly
- The change in energy can be considered as a sum of heat and work done on and by the system.  
 $\Delta E$  is a state function,  $q$  and  $w$  are not.

$$\boxed{\Delta E = q + w}$$

**$\Delta E$**  is the change in Energy

**$q$**  is heat (in J)

**$w$**  is work (in J)

#### WIZE TIP

Some profs write  **$\Delta E$**  and others write  **$\Delta U$** . They mean the same thing! Both are referring to the internal energy of the system!

From the 1<sup>st</sup> law of thermodynamics we can see that change in energy for system and surroundings must be equal in magnitude because the **change in internal energy of the universe must be equal to zero.**



$$\Delta E_{univ} = \Delta E_{sys} + \Delta E_{surr} = 0$$

$$\Delta E_{surr} = -\Delta E_{sys}$$

#### Examples:

- If the system gains 5 J of energy, the surroundings (gains/loses) loses 5 J of energy.
- If the system loses 10 J of energy, the surroundings (gains/loses) gains 10 J of energy.

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## Internal Energy of Ideal Gases

The internal energy of an ideal gas is only dependent on temperature:

$$\Delta E(T)$$

So for any Isothermal process involving an ideal gas,

$$\Delta E = 0$$

## Signs of Heat and Work & Calculating Work

Recall the equation for the first law of thermodynamics:

$$\Delta E = q + w$$

**ΔE** is the change in Energy

**q** is heat (in J)

**w** is work (in J)

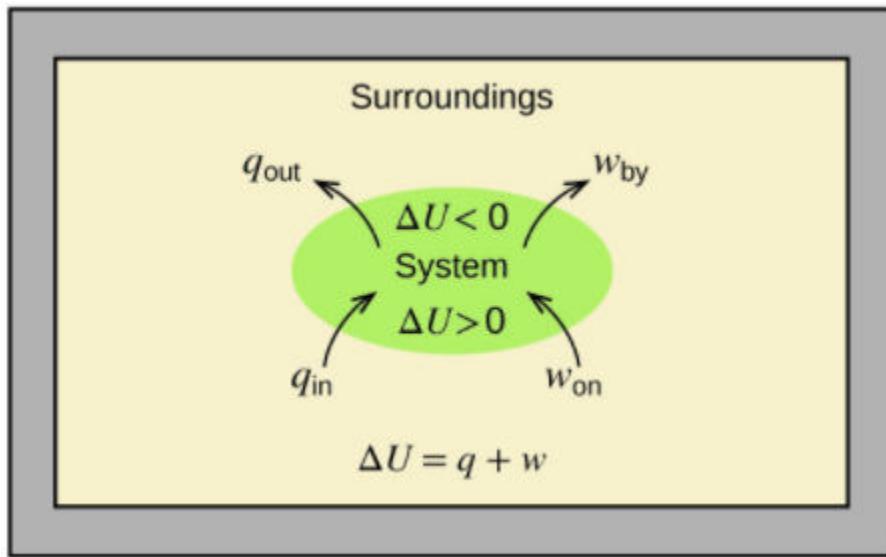


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In chemistry **everything is with respect to the system** so the following conventions exist:

### WIZE CONCEPT

**w > 0** if the surroundings do work on the system

**w < 0** if the system does work on the surroundings

**q > 0** if heat moves from the surroundings to the system

**q < 0** if heat moves from the system to the surroundings

---

When the external pressure is constant we can **calculate work** using the following equation:

$$w = -P_{ex}\Delta V$$

$w$ =work (J)

$P_{ext}$ =constant external pressure (in kPa)

$\Delta V=V_f-V_i$  (in L)

## Example: Work Heat and Internal Energy of an Ideal Gas

20 L of an ideal gas isothermally expands to 40 L against a constant external pressure of 1 atm. Calculate  $w$ ,  $q$ , and  $\Delta E$  for this process.

$$W = -P_{ex} (V_2 - V_1)$$

$$P_{ex} = 1 \text{ atm} = 101,325 \text{ kPa} \quad V_2 = 40 \text{ L} \quad V_1 = 20 \text{ L}$$

$$\begin{aligned} W &= -(101,325)(40 - 20) \\ &= -2026.5 \text{ J} \end{aligned}$$

$$q = -w = 2026.5 \text{ J}$$

---

**3.4.4**

A system composed of 4.50 L of N<sub>2</sub> in a cylinder expands against an external pressure of 300 kPa until its volume is 6.30 L.

- a. What is the value of w for this process?
- b. What would the value of w be if the external pressure was 0 kPa?

a)

$$-540 \text{ J}$$

b)

$$0$$

a)  $W = -P_{\text{ext}} \Delta V \quad P_{\text{ext}} = 300 \text{ kPa} \quad \Delta V = V_2 - V_1 = 6.30 - 4.50 = 1.80 \text{ L}$   
 $= -(300)(1.8) = -540 \text{ J}$

b)  $W = -P_{\text{ext}} \Delta V \quad P_{\text{ext}} = 0$   
 $= 0$

## 3.4.5

Fill in the following blanks!

How to fill in this table:

For the first column we have  $\Delta U$ . Your options are temperature is constant, temperature increases, or temperature decreases

For the second column we have  $q$ . Your options are no heat is exchanged, heat enters the gas, or heat exits the gas

For the third column we have  $w$ . Your options are the volume is constant, the gas is compressed, or the gas expands.

	$\Delta U$ (Change in Internal Energy):	$q$ (heat):	$w$ (work):
Is positive when...	T increase	heat absorbed by system	gas compressed, volume ↓
Is negative when...	T decrease	heat loss from system	gas expands, volume ↑
Is 0 when...	$\Delta T = 0$	adiabatic process no heat added or removed	constant volume

## 3.5 Heating Curves

3.5.1

### Introduction to the Heat Curve

Before we take a look at what a heat curve is, we must first understand what can happen when heat is added to a system.



#### WIZE CONCEPT

When heat is added to a system, the system can EITHER:

1) Experience a phase change:

$$q = n\Delta H_{\text{phase change}}$$

or depending on units it could be

$$q = m\Delta H_{\text{phase change}}$$

OR

2) Experience a temperature change:  $q = mc\Delta T$

---

## More Information on Each of These Important Equations

$$q = n\Delta H_{phase\ change}$$

***q is heat***

***n is moles***

**$\Delta H_{phase\ change}$**  is the enthalpy value that describes the **heat of transition** (energy required to complete a transition)

**Example:** the heat of vapourization ( $\Delta H_{vap}$ ) is describing the amount of energy that is required to convert a liquid → gas (vapourization).

$$q = mc\Delta T$$

*q is heat*

*m is mass (g)*

*c is the specific heat capacity (J/g°C or J/g°K)*

*ΔT is the change in temperature ΔT=Tf-Ti (°C or °K)*

 **WIZE TIP**

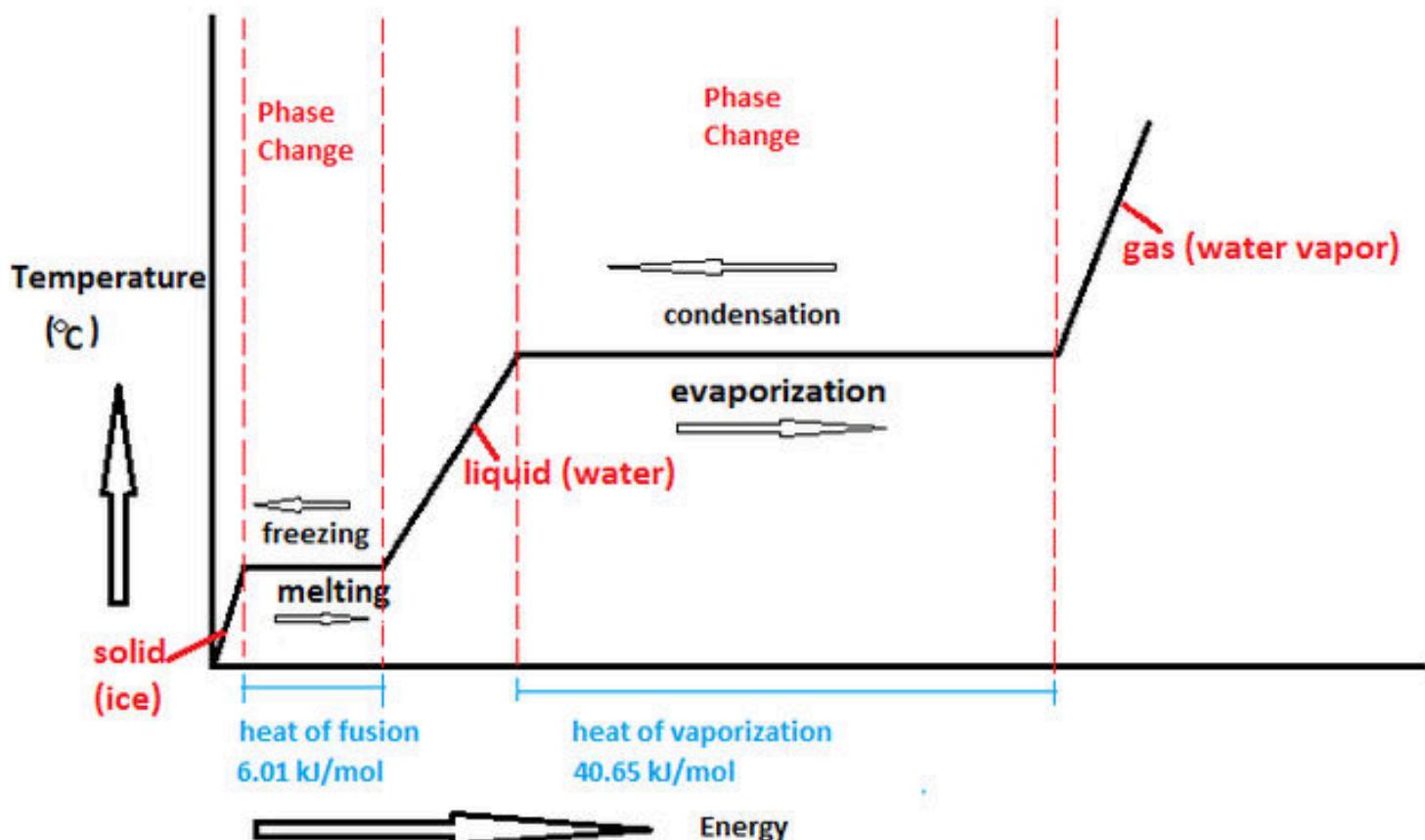
For H<sub>2</sub>O(l), c=4.18J/g°C. This means that it takes 4.2 J of heat energy to heat up 1 g of liquid water by 1°C!

**Note:** Each phase will have a different "c" value!

**Example:** We will see that water in liquid form and solid ice form have different c values!

- H<sub>2</sub>O(l), c=4.18 J/g°C
- H<sub>2</sub>O(s), c=2.06 J/g°C

## Heat curve for H<sub>2</sub>O



- Label the melting and boiling points of water on the y axis.
- Label what is happening on the horizontal lines as we move to the right.
- Label the appropriate equation for each column in the diagram ( $q=mc\Delta T$  or  $q=n\Delta H$ ).
  - What  $c$  values and  $\Delta H$  values would you use in each column?
- What is the slope of horizontal sections? 0
- Are the phase changes occurring at the parts of the graph that have a slope of 0 or a positive slope? 0

## Summary

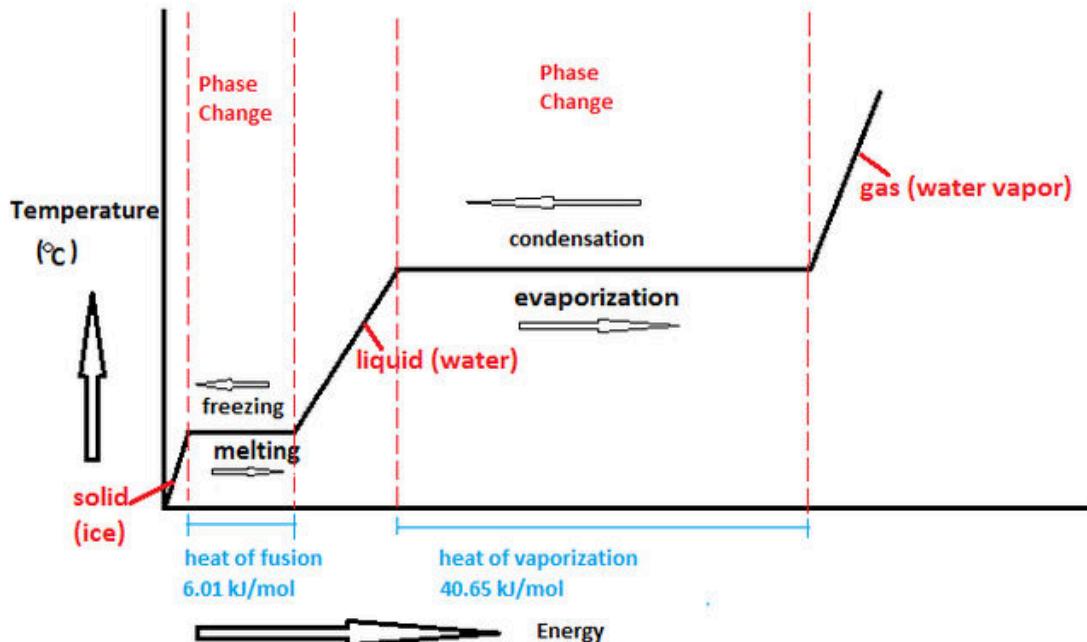


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### WIZE CONCEPT

The horizontal regions on the graph show when a phase change is occurring. We know this because there is no  $\Delta T$  at these horizontal lines.

We said we can either have a phase change OR a temperature change.

During a phase change, all the heat is being used to overcome the intermolecular forces to make the phase change possible!

Once a phase change has occurred, then the heat increases the temperature (this is where we see the positive slopes)!

## Important Things to Know About the Heat Curve

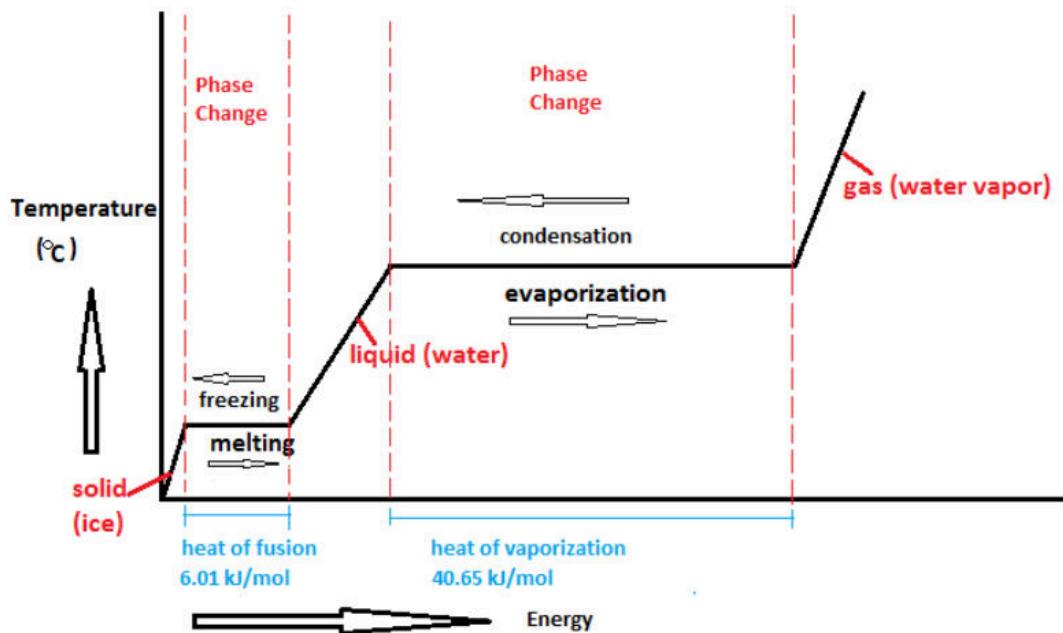


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Do you think the  $\Delta H_{\text{fusion}}$  or  $\Delta H_{\text{vapourization}}$  would be a larger number? Why? (You can look at the heat curve to help you decide)

$\Delta H_{\text{vap}}$  would be a larger number since there are more intermolecular forces that needed to be break to go from liquid to gas than from solid to gas

- This will always be the case!
- The longer the horizontal line in the heat curve, the more/less: more the value of the  $\Delta H$  for that phase change!
  - $q_{\text{vap}}$  will also be greater than  $q_{\text{fusion}}$

Based on the given values of  $\Delta H_{\text{fusion}}$  and  $\Delta H_{\text{vapourization}}$  in the diagram above, what do you think the  $\Delta H_{\text{crystallization}}$  and  $\Delta H_{\text{condensation}}$  would be?

Same magnitude but differ sign.

$$\Delta H_{\text{crys}} = -6 \text{ kJ/mol} \quad \Delta H_{\text{conden}} = -40.65 \text{ kJ}$$

## Melting/Boiling Points

The **melting point** is the temperature where (s)  $\rightleftharpoons$  (l) phase

The **boiling point** is the temperature where (l)  $\rightleftharpoons$  (g) phase

Both these points (melting and boiling) occur at a pressure of 1atm. We'll see how this will come in handy for us to know soon!

## Example: Heat Curve

What is the amount of heat required to melt 10 g of ice at -5°C and convert it to liquid at 60°C?

$$c_{\text{ice}} = 2.06 \text{ J/g}^{\circ}\text{C}, \Delta H_{\text{fus}} = 334 \text{ J/g}, c_{\text{H}_2\text{O(l)}} = 4.18 \text{ J/g}^{\circ}\text{C}$$

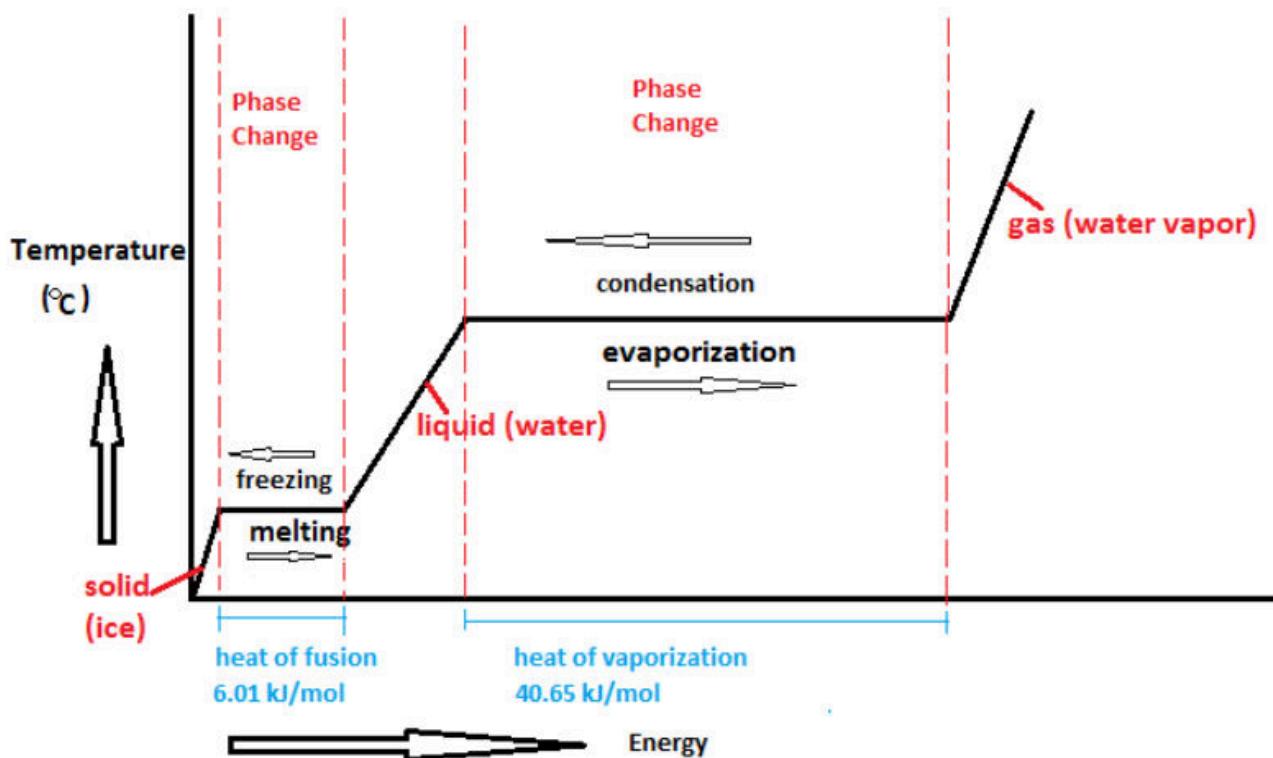


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- state
- 1) at ice  $-5^{\circ}\text{C} \rightarrow 0^{\circ}\text{C}$
  - 2) melting
  - 3) water  $0^{\circ}\text{C} \rightarrow 60^{\circ}\text{C}$

$$\begin{aligned}
 q_T &= m C_{\text{ice}} \Delta T + \Delta H_{\text{fusion}} + m C_{\text{water}} \Delta T \\
 &= (10 \text{ g})(2.06 \text{ J/g}^{\circ}\text{C})(5^{\circ}\text{C}) + (10 \text{ g})(334 \text{ J/g}) + (10 \text{ g})(4.18 \text{ J/g}^{\circ}\text{C}) \\
 &= 5951 \text{ J}
 \end{aligned}$$

**Hint:** For a problem like this we need to think of what we are doing (increasing temperature or changing phase) and when

1. We start with ice at  $-5^{\circ}\text{C}$ . The ice will need to be heated until it reaches  $0^{\circ}\text{C}$
2. At  $0^{\circ}\text{C}$  we then change ice's phase into liquid by melting it
3. Once we melt the ice into a liquid, we then heat the liquid from  $0^{\circ}\text{C}$  to  $60^{\circ}\text{C}$

**Now that we have thought of the different steps, let's write out an equation that describes each step:**

**If we wanted to answer in kJ, what would be our answer (if our answer only has 2 significant digits):**

**! WATCH OUT!**

Pay attention to the units for these problems, convert units if necessary.

3.5.4

## Practice: Solve for Heat

If we are told that water's heat of fusion ( $\Delta H_{\text{fus}}$ ) is 6 kJ/mol, then how much heat would be required to melt 5 moles of water?

-6 kJ

6 kJ

25 kJ

30 kJ



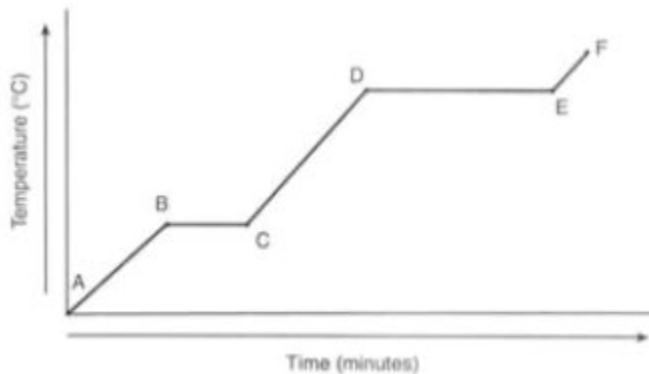
$$\Delta H_{\text{fus}} = 6 \text{ kJ/mol} \quad n = 5 \text{ mol}$$

$$\begin{aligned} q &= n \Delta H_{\text{fus}} \\ &= (5 \text{ mol})(6 \text{ kJ/mol}) \\ &= 30 \text{ kJ} \end{aligned}$$

3.5.5

## Practice: Heat Curve

Consider the heat curve below and the following statements:



- 1 The segment A-B corresponds to the heating of a solid.
2. The segment D-E corresponds to the vaporization of a liquid.
3. The segment E-F corresponds to the heating of a liquid
4. The segment B-C corresponds to a phase change.

Only 1 is false.

Only 2 is false.

Only 3 is false.

All of them are false.

Only 2, 3, and 4 are false.

## 3.6 Calorimetry

3.6.1

### Heat Capacity and Calorimetry

- **Heat capacity (C):** has units of  $J/^{\circ}K$  or  $J/^{\circ}C$  and describes the amount of heat needed to heat up the substance or object by one degree
- **Specific heat capacity (c):** has units of  $J/g^{\circ}K$  or  $J/g^{\circ}C$  and described the amount of heat required to heat up 1g of a substance by one degree.
- We also have a **molar heat capacity** (units  $J/Kmol$ , or  $J/Cmol$ ), which is the amount of heat required to heat up 1 mol of a substance by one degree).

$$q = C (\Delta T) \text{ or } q = mc (\Delta T) \text{ or } q = nC (\Delta T)$$

**q** is heat (in J)

**C** is the heat capacity ( $J/^{\circ}K$  or  $J/^{\circ}C$ )

**c** is the specific heat capacity ( $J/g^{\circ}K$  or  $J/g^{\circ}C$ )

**m**=mass (in g)

**ΔT**=change in temperature ( $T_2-T_1$ ) in  $^{\circ}K$  or  $^{\circ}C$

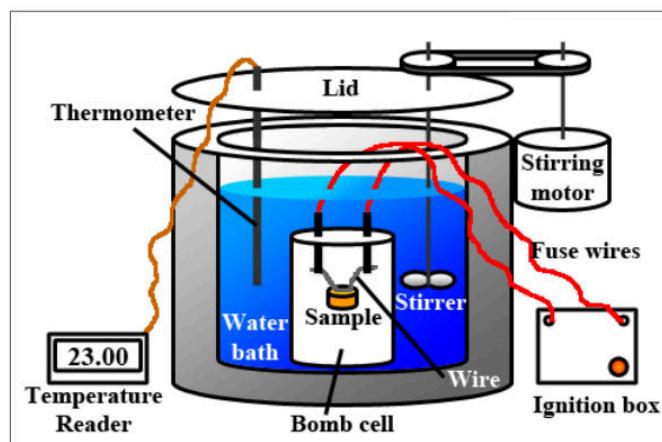
**n**=moles

**Note:**  $mc=C$

## General Introduction to Calorimeters

Calorimeters come in two general varieties:

- Simple "coffee cup" calorimeters
- And bomb calorimeters



Left Photo by Community College Consortium for Bioscience Credentials / CC BY, right photo by Lisdavid89 / CC BY

What is a calorimeter?

- A calorimeter is just a container that is insulated. It holds a liquid that is usually water inside of it and you can have reactions happening inside of it as well.
- For a calorimeter, the **water is the surroundings and the reaction is the system**.
- A thermometer will be able to read the temperature changes of the water.
  - If the thermometer measures an increase in temperature, that means the water went up in temperature.
  - If the water went up in temperature then that means it gained all of the heat from the system (or the reaction taking place inside the calorimeter). It didn't gain the heat from anywhere else because the calorimeter is insulated!
  - This means the system/reaction lost heat and is endothermic/exothermic: exothermic



WIZE CONCEPT

In general:

$$q_{\text{water}} = -q_{\text{rxn}} \text{ where } q = mc\Delta T$$

## Coffee Cup Calorimeters

Essentially it's an open styrofoam cup with a thermometer.



These are most frequently used to measure the temperature change of an aqueous reaction under constant pressure conditions.

$$\Delta H = \Delta E + P\Delta V$$

$$\Delta H = q + w - w$$

$$\Delta H = q_p \text{ (under constant pressure conditions)}$$

### WIZE CONCEPT

In general:

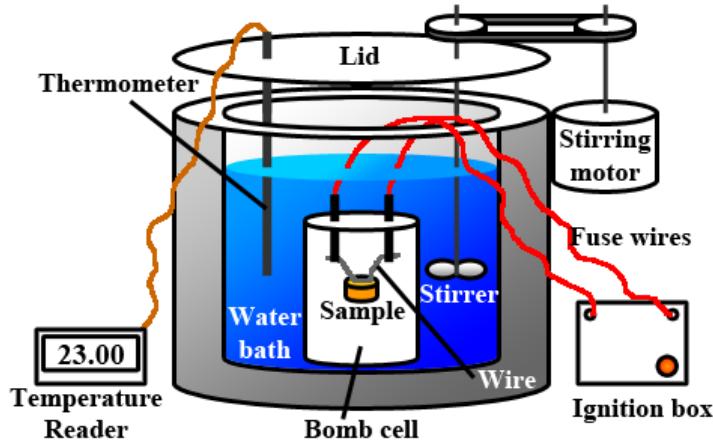
$$q_{\text{system}}(\text{sample}) = -q_{\text{surroundings}}(\text{H}_2\text{O})$$

using  $q = mc\Delta T$ :

$$(\text{mass of sample})(c \text{ of sample})(\Delta T \text{ of sample}) = -(\text{mass of H}_2\text{O})(c \text{ of H}_2\text{O})(\Delta T \text{ of water})$$

## Bomb Calorimeters

Typically these consist of a sealed reaction vessel loaded with a combustible material and filled with oxygen placed in a surrounding water bath.



A combustion reaction is performed under constant volume conditions and the resulting change in temperature of the water bath is measured.

$$\Delta E = q_v + w$$

$$\Delta E = q_v - P\Delta V$$

$(\Delta V=0)$  so  $w=0$

$$\Delta E = q_v$$

For bomb calorimeters, we will see two types of problems:

- 1) Sometimes we evaluate the heat capacity for the entire calorimeter (the  $q$  of the calorimeter represents bomb and the water bath)
- 2) Sometimes we evaluate the heat capacity of the bomb and water separately ( $q_{\text{total}} = q_{\text{bomb}} + q_{\text{water}}$ )

**i WIZE TIP**

Read the question carefully to figure out which one of these types of problems we are looking at!

## Example: Copper Weight Calorimetry

A copper weight (52.3 g) is heated by a flame from room temperature ( $25^\circ C$ ) to a temperature of  $189^\circ C$ . The specific heat capacity of copper is  $0.385 \text{ J g}^{-1} \text{ K}^{-1}$

a) What is  $\Delta H$  for this process?

$$m = 52.3 \text{ g} \quad C_p = 0.385 \text{ J/g}\cdot\text{K} \quad \Delta T = 189 - 25 = 164$$

$$\begin{aligned} \Delta H = q_p &= m C_p \Delta T = (52.3 \text{ g})(0.385 \text{ J/g}\cdot\text{K})(164 \text{ K}) \\ &= 33027 \text{ J} \end{aligned}$$

b) The block is then allowed to cool in some water until the temperature of the block is  $75^\circ C$ , what is  $\Delta H$  for this process?

$$\begin{aligned}\Delta H = q_p &= m C_p \Delta T = (52.3\text{g})(0.385\text{J/g}\cdot\text{K})(75 - 189)\text{K} \\ &= -2295 \text{ J}\end{aligned}$$

c) What is the  $\Delta H$  from part b) with respect to the water ?

$$q_{\text{water}} = -q_{\text{cu}} = +2295 \text{ J}$$

## Example: Calorimetry

A block of 20 g lead cube ( $c_p = 0.128 \text{ J g}^{-1}\text{K}^{-1}$ ) is heated to  $145^\circ \text{ C}$ . The cube is then submerged into a Styrofoam cup containing 500 g of water at room temperature ( $c_p = 4.18 \text{ J g}^{-1}\text{K}^{-1}$ ). After the water and the lead cube come to thermal equilibrium, what is the final temperature of the water?

$$q_{\text{lead}} = -q_{\text{water}}$$

$$m c_p \Delta T = -m c_p \Delta T$$

$$(20)(0.128)(T_2 - 418) = -500(4.18)(T_2 - 298)$$

$$2.56T_2 - 1070.80 = -2090T_2 + 622820$$

$$(2.56 + 2090)T_2 = 622820 + 1070.80$$

$$T_2 = 298.15 \text{ K}$$

## 3.6.4

The temperature of a 12.58 g sample of calcium carbonate ( $\text{CaCO}_3(s)$ ) increases from  $23.6^\circ\text{C}$  to  $38.2^\circ\text{C}$ . If the specific heat capacity of calcium carbonate is 0.82 J/g K. How many joules of heat are absorbed?

151 J



5.0 J



7.5 J



410 J



0.82 J



$$q = mc\Delta T.$$

$$M = 12.58 \text{ g} \quad c = 0.82 \text{ J/g}\cdot\text{K} \quad T_i = 23.6^\circ\text{C} \quad T_f = 38.2^\circ\text{C}$$
$$\Delta T = 14.6^\circ\text{C}$$

$$q = 12.58(0.82)(14.6)$$
$$\approx 152.8 \text{ J}$$

## 3.6.5

The  $\Delta H_{\text{soln}}$  for the process when solid sodium hydroxide (NaOH) dissolves in water is 44 kJ/mol. When a 10.0 g sample of NaOH dissolves in 250.0 g of water in a coffee-cup calorimeter, the temperature increases from 23.0°C to \_\_\_\_°C. Assume that the solution has the same specific heat as liquid water, i.e. 4.18 J/gK.

35.2



24.0



33.5



40.2



$$q = mc\Delta T$$

$$m = 250 \text{ g} \quad c = 4.18 \text{ J/gK}$$

$$\Delta H_{\text{sol}} = 44 \text{ kJ/mol} \quad M_{\text{sample}} = 10.0 \text{ g} \quad M_{\text{NaOH}} = 40.00 \text{ g/mol}$$

$$n_{\text{sample}} = \frac{10 \text{ g}}{40 \text{ g/mol}} = 0.25 \text{ mol}$$

$$q_{\text{NaOH}} = 0.25 \text{ mol} \times 44 \text{ kJ/mol} = 11 \text{ kJ}$$

$$\Delta T = \frac{11,000}{250 \times 4.18} \approx 10.53 \text{ K}$$

$$T_f = T_i + \Delta T = 23^\circ\text{C} + 10.53^\circ\text{C} = 33.5^\circ\text{C}$$

## Example: Bomb Calorimetry

When 2g of rocket fuel,  $\text{N}_2\text{H}_4$  is burned inside of a bomb-calorimeter, the temperature of the water rises from  $22^\circ\text{C}$  to  $29^\circ\text{C}$ . If there is 0.9kg of water and the heat capacity for the bomb is  $822\text{J}/^\circ\text{C}$ , what is the heat of the combustion for one mole of  $\text{N}_2\text{H}_4$  in the bomb calorimeter?

- a) 5000 J
- b) -51 000 J
- c) -513 kJ
- d) -573 kJ

$$m_{\text{N}_2\text{H}_4} = 2\text{ g}$$

$$T_1 = 22^\circ\text{C} \quad T_2 = 29^\circ\text{C} \quad \Delta T = 29 - 22 = 7^\circ\text{C}$$

$$m_{\text{H}_2\text{O}} = 0.9\text{ kg} = 900\text{ g}$$

$$C_{(\text{bomb})} = 822\text{ J}/^\circ\text{C}$$

$$\begin{aligned} q_{\text{bomb}} &= C_{\text{bomb}} \Delta T \\ &= 822\text{ J}/^\circ\text{C} \times (7^\circ\text{C}) = 5754\text{ J} \end{aligned}$$

$$q_{\text{water}} = m C \Delta T = 900 (4.18)(7) = 26334\text{ J}$$

$$q_T = 5754\text{ J} + 26334\text{ J} = 32088\text{ J}$$

$$q_{\text{cal}} = 32088\text{ J} \quad 0.0625\text{ mol } \text{N}_2\text{H}_4 \text{ released } 32088\text{ J}$$

$$M_{\text{N}_2\text{H}_4} = 32\text{ g/mol}$$

$$n = \frac{2}{32} = 0.0625\text{ mol} \Rightarrow$$

1 mol  $\text{N}_2\text{H}_4$  will release  $513408\text{ J}$

$\hookrightarrow -513\text{ kJ}$

## Practice: Bomb Calorimetry

A fuel compound was burned in oxygen in a bomb calorimeter with a combined heat capacity of 56.5 kJ/K (Bomb assembly and water). The temperature of the surrounding water bath increased from 23.5 to 34.6 °C as a result. What was the value of  $q_{sys}$  for this process?

A) +333 kJ

B) -333 kJ

C) +627 kJ

D) -627 kJ

$$q_{sys} = C \Delta T$$

$$C = 56.5 \text{ kJ/K} \quad T_f = 34.6^\circ\text{C} \quad T_i = 23.5^\circ\text{C} \quad \Delta T = 11.1 \text{ K}$$

$$\begin{aligned} q_{sys} &= 56.5 \times 11.5 \\ &= 627.15 \text{ kJ} \end{aligned}$$

$\therefore$  since the rxn is combustion process so heat is released by the system,  $q_{sys} = -627.15 \text{ kJ}$ .

## 3.7 Heats of Formation Method

3.7.1

### Heats of Formation

**Heat of formation** is the amount of heat that is required to **form 1 mole of a compound from its constituent elements in their natural/standard state** (the state they are in under standard conditions):

$$P = \frac{1}{101325} \text{ atm}, T = \frac{25}{298} \text{ }^{\circ}\text{C}, \frac{1}{1} \text{ M concentration, pH } \frac{7}{7}$$
$$101325 \text{ Pa} = 101.325 \text{ kPa}$$

#### WIZE CONCEPT

The  $\Delta H^{\circ f}$ =0 for an element in standard state.

#### WIZE TIP

\*\*The following elements in their standard state should be memorized and the phase they are in is important too!

C(s) as graphite is in standard state.

I<sub>2</sub>(s)

Br<sub>2</sub>(l)

Hg(l)

Diatomc molecules (BrINCIHOF)

*Examples:*

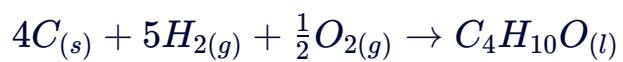
Cl<sub>2</sub>(g), H<sub>2</sub>(g), O<sub>2</sub>(g)

#### WATCH OUT!

If you are shown O(g) on your exam, this is NOT oxygen in its standard state. Remember oxygen in its standard state has to be gaseous AND diatomic! O<sub>2</sub>(g) would be in standard state.

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*Example:* The **formation reaction** for C<sub>4</sub>H<sub>10</sub>O<sub>(l)</sub> would be:



The enthalpy of formation for this compound is -327kJ/mol.

## Calculating the Enthalpy of Reaction Using Heats of Formations

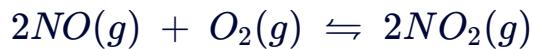
If you know the enthalpies of formation for each reactant and product in any chemical equation, you can find the enthalpy change for that reaction:

$$\Delta H^o_{rxn} = \left[ \left( \sum n \Delta H^o f_{products} \right) - \left( \sum n \Delta H^o f_{reactants} \right) \right]$$

In other words, **Δ= Final - Initial**

## Example: Calculating the Enthalpy of a Reaction

Calculate  $\Delta H^\circ_{rxn}$  for the following reaction:



$$\Delta H^\circ_f(NO(g)) = 90.75 \text{ kJ/mol}$$

$$\Delta H^\circ_f(NO_2(g)) = 33.18 \text{ kJ/mol}$$

$$\underline{\Delta H^\circ_f(O_2(g)) = 0 \text{ kJ/mol}}$$

because  $O_2$  is already in standard state which oxygen naturally exists under standard conditions, so enthalpy of formation of  $O_2$  is zero

$$\begin{aligned} \sum \Delta H_f^\circ(\text{products}) &= 2 \sum \Delta H_f^\circ(NO_2) \\ &= 2 \times 33.18 = 66.36 \text{ kJ} \end{aligned}$$

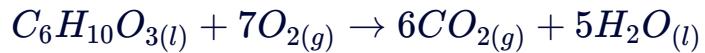
$$\begin{aligned} \sum \Delta H_f^\circ(\text{reactants}) &= 2 \sum \Delta H_f^\circ(NO) + 1 \sum \Delta H_f^\circ(O_2) \\ &= 2(90.75) + 1(0) = 181.5 \text{ kJ} \end{aligned}$$

$$\begin{aligned} \Delta H_{rxn}^\circ &= 66.36 - 181.5 \\ &= -115.14 \text{ kJ} \end{aligned}$$

$\therefore \Delta H_{rxn}^\circ$  is negative, the process is exothermic.

## Practice: Calculate the Enthalpy of the Reaction

Find  $\Delta H^\circ_{rxn}$  for the following reaction at 25°C:



Given the following data:

Species	$\Delta^\circ H_f$
$C_6H_{10}O_{3(l)}$	-640.4 kJ/mol
$CO_{2(g)}$	-393.5 kJ/mol
$H_2O_{(l)}$	-285.8 kJ/mol

A) -4220 kJ

B) -3150 kJ

C) -2250 kJ

D) +1220 kJ

$$\begin{aligned}\sum H_f^\circ(\text{Product}) &= 6\sum H_f^\circ(CO_2) + 5 \sum H_f^\circ(H_2O) \\ &= 6 \times (-393.5) + 5 \times (-285.8) \\ &= -2361 + (-1429) = -3790 \text{ kJ}\end{aligned}$$

$$\begin{aligned}\sum H_f^\circ(\text{Reactant}) &= 1 \sum H_f^\circ(C_6H_{10}O_3) + 7 \sum H_f^\circ(O_2) \\ &= -640.4 + 7 \times 0 = -640.4 \text{ kJ}\end{aligned}$$

$$\begin{aligned}\sum H_{rxn}^\circ &= \sum H_f^\circ(\text{Product}) - \sum H_f^\circ(\text{Reactant}) \\ &= -3790 - (-640.4) = -3149.6 \text{ kJ}\end{aligned}$$