أوراق عمل الوحدة الأولى Change Chemical and Energy الطاقة والتغيرات الكيميائية

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12 Advanced Chemistry 2023

Chapter 1

Energy and Chemical Change

Section 1: Energy
Section 2: Heat
Section 3: Thermochemical Equations
Section 4: Calculating Enthalpy Change
Section 5: Reaction Spontaneity

Manasra Academy

Work Sheets In Chemistry

CHAPTER 1 / Energy and Chemical Changes

Grade

12 Advance / T1

By

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NAME :………………………………………………….

CLASS : ………………………………………………….

SCHOOL: ……………………………………………….

CHAPTER 1 / Energy and Chemical Changes

Section 1 : Energy

- **Energy:** is the ability to do work or produce heat.
- It exist in tow basic forms:
  - potential energy
  - kinetic energy.
- **Potential energy is energy due to the composition or position of an object.**
  An example:downhill skier poised at the starting gate for a race (as shown in figure 1a)
- **Kinetic energy : is energy of motion.**
  An example: the potential energy of skier changes to kinetic energy during the speedy trip to the finish line (as shown in figure 1b)

Figure 1 At the top of the course, the skier in **a** has high potential energy because of her position. In **b,** the skier's potential energy changes to kinetic energy.

Compare *How is the potential energy of the skier different at the starting gate and at the finish line?*

- Chemical systems contain both kinetic energy and potential energy.
- The kinetic energy of a substance :
  - is directly related to the constant random motion of its particles
  - and is proportional to temperature , as temperature increase the motion of particles increase .
- The potential energy of a substance depends on its composition:
  - The tybe of atoms in the substance
  - The number and type of chemical bonds joining the atoms
  - The arrangement of the atoms.

Law of conservation of energy:

- Energy changes from one form to another, but energy is conserved.the total amount of energy remains constant.
- **The law of conservation of energy (or The first law of thermodynamics)** states that: in any chemical reaction or physical process, energy can be converted from one form to another , but it is neither created nor destroyed.
- **Question :**

The process	The change of energy	
	From	To
1- Water rushes through turbines in the hydroelectric plant.		
2- Burns of propane gas (used in cooking and heating)		
3- Burns of gasoline (octane) in the engine of the car.		
4- $27\ kJ + NH_4NO_{3(s)} \rightarrow NH^+_{4\,(aq)} + NO^-_{3\,(aq)}$		
5- $4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)} + 1625\ kJ$		
6- avalanche An of snow		
7- A stock-car race		

Chemical potential energy:

- It is the energy that is stored in a substance because of its composition.

Heat (q):is energy that is in the process of flowing from a warmer object to a cooler object.

Measuring Heat:

- **Calorie (cal): the amount of energy required to raise the temperature of 1g of pure water by 1C° .**
- **Nutritional calorie = Cal**
- **1Cal = 1000cal (kcal)**
- **The SI unit of energy and of heat is the joul.**

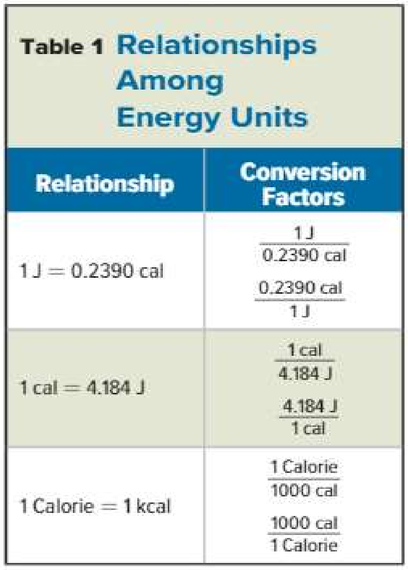
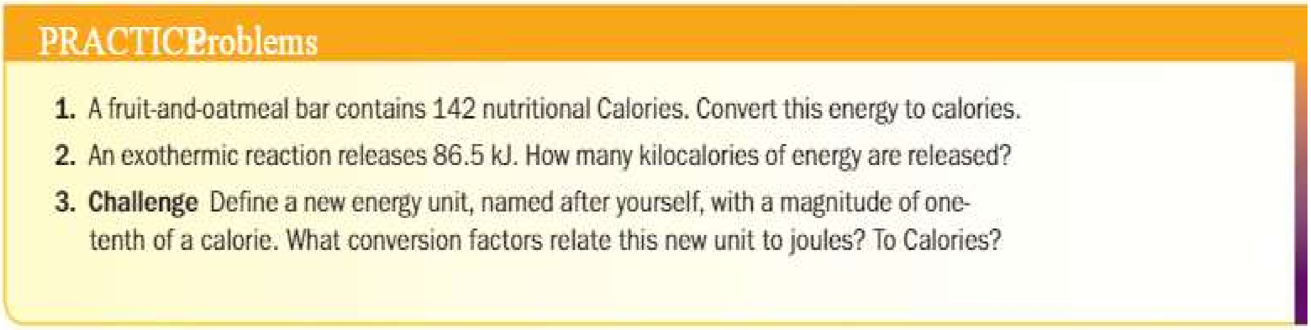


Table 1 Relationships Among Energy Units

Relationship	Conversion Factors
1 J = 0.2390 cal	$\frac{1\ J}{0.2390\ cal}$ $\frac{0.2390\ cal}{1\ J}$
1 cal = 4.184 J	$\frac{1\ cal}{4.184\ J}$ $\frac{4.184\ J}{1\ cal}$
1 Calorie = 1 kcal	$\frac{1\ Calorie}{1000\ cal}$ $\frac{1000\ cal}{1\ Calorie}$



PRACTICE Problems

1. A fruit-and-oatmeal bar contains 142 nutritional Calories. Convert this energy to calories.
2. An exothermic reaction releases 86.5 kJ. How many kilocalories of energy are released?
3. **Challenge** Define a new energy unit, named after yourself, with a magnitude of one-tenth of a calorie. What conversion factors relate this new unit to joules? To Calories?

Specific Heat:

- It is the amount of heat required to raise the temperature of 1g of that substance by $1C^{\circ}$
- Each substance has its own specific heat **because** different substances have different composition.
- For example water has specific heat of = $4.184 J/g.C^{\circ}$. The specific heat of concrete is = $0.84 J/g.C^{\circ}$.
- This means that the specic heat of water is five times of concrete.
- When equal masses of concrete and water absorb equal amounts of energy , the temperature of concrete increases five times of water

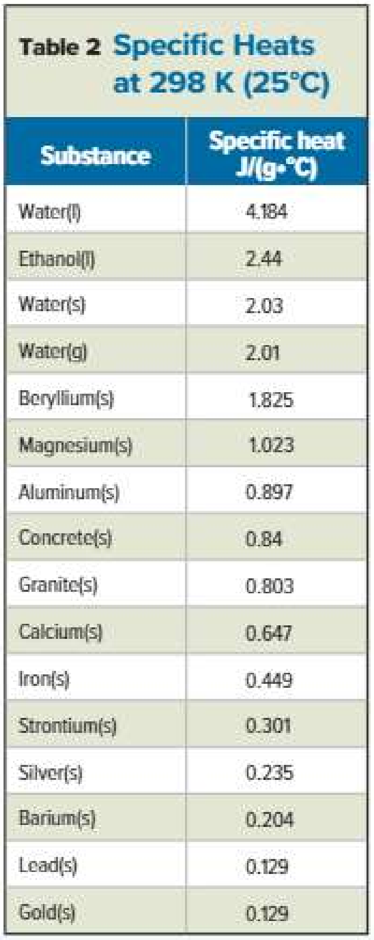
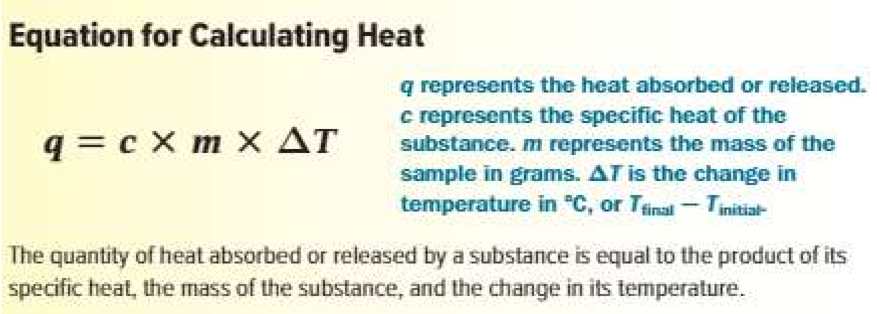


Table 2 Specific Heats at 298 K (25°C)

Substance	Specific heat J/(g·°C)
Water(l)	4.184
Ethanol(l)	2.44
Water(s)	2.03
Water(g)	2.01
Beryllium(s)	1.825
Magnesium(s)	1.023
Aluminum(s)	0.897
Concrete(s)	0.84
Granite(s)	0.803
Calcium(s)	0.647
Iron(s)	0.449
Strontium(s)	0.301
Silver(s)	0.235
Barium(s)	0.204
Lead(s)	0.129
Gold(s)	0.129

Calclating heat absorbed and heat released:



Equation for Calculating Heat

$$q = c \times m \times \Delta T$$

q represents the heat absorbed or released. *c* represents the specific heat of the substance. *m* represents the mass of the sample in grams. ΔT is the change in temperature in °C, or $T_{final} - T_{initial}$.

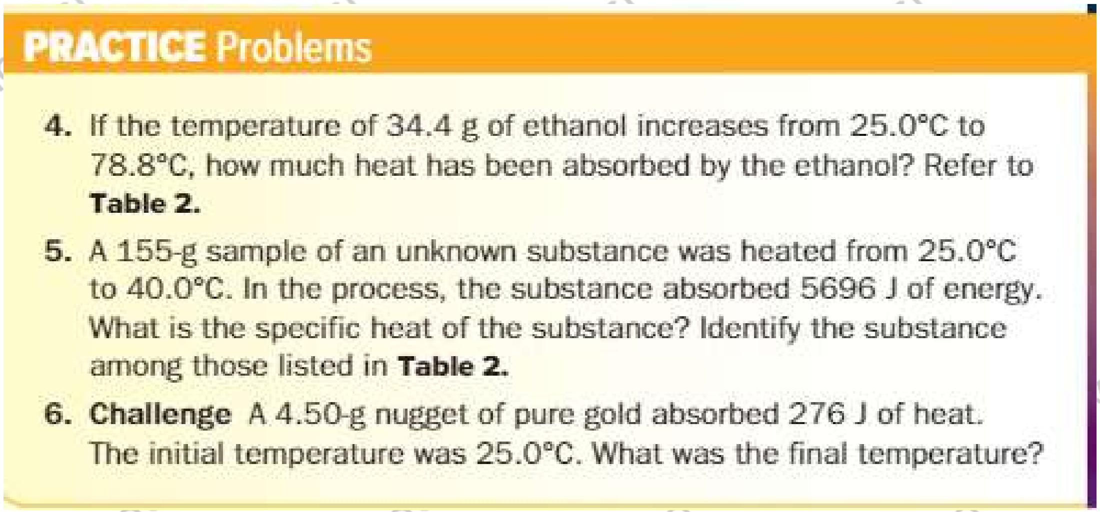
The quantity of heat absorbed or released by a substance is equal to the product of its specific heat, the mass of the substance, and the change in its temperature.

- **Example 1** :calculate the heat absorbed by equal masses of concrete and water ($5x10^3$g) When its temperature increased by 6C°, and compare between them .

- Example 2 :Calculate the energy released by $5x10^3$g of concrete reached a temperature of 74.0C° during asunny day and cooled down to 40.0C° at night.

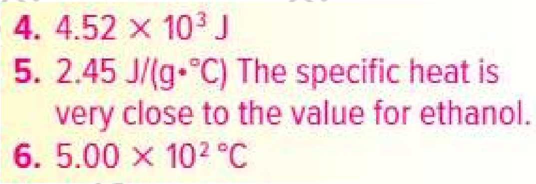
- Equal masses of aluminum, gold, iron, and silver were left to sit in the Sun at the same time and for the same length of time , arrange the four metals according to the increase in their temperatures from largest increase to smallest.

The element	Al	Au	Fe	Ag
Specific Heat J/g.C°	0.897	0.129	0.449	0.235



PRACTICE Problems

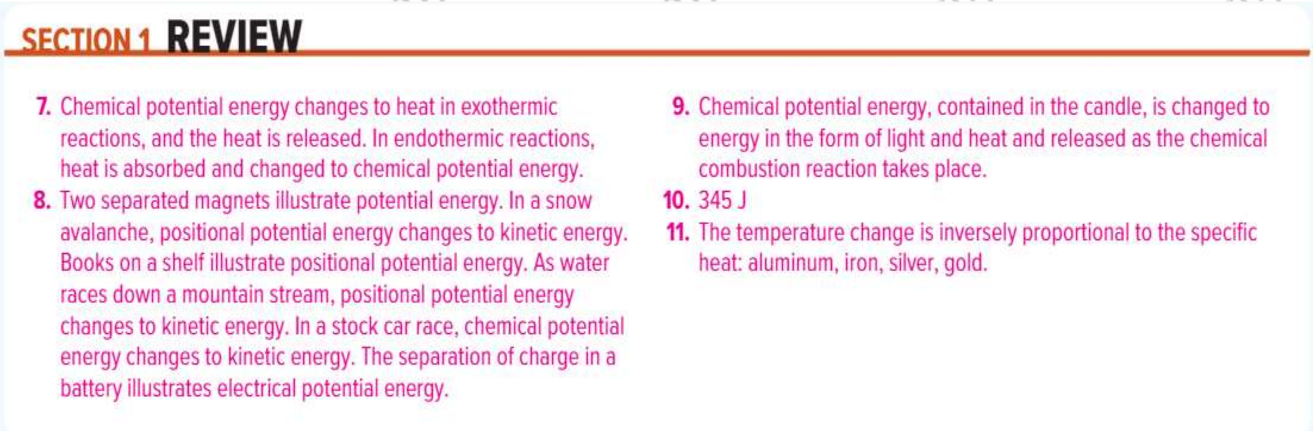
4. If the temperature of 34.4 g of ethanol increases from 25.0°C to 78.8°C, how much heat has been absorbed by the ethanol? Refer to **Table 2.**
5. A 155-g sample of an unknown substance was heated from 25.0°C to 40.0°C. In the process, the substance absorbed 5696 J of energy. What is the specific heat of the substance? Identify the substance among those listed in **Table 2.**
6. **Challenge** A 4.50-g nugget of pure gold absorbed 276 J of heat. The initial temperature was 25.0°C. What was the final temperature?



4. 4.52×10^3 J
5. 2.45 J/(g•°C) The specific heat is very close to the value for ethanol.
6. 5.00×10^2 °C

Using the sun's energy:

- Water is sometimes used to harness the energy of the sun **because** of its high specific heat to provide heat in homes.
- Radiation from the Sun could supply all the energy needs of the world and reduce the use of carbon dioxide producing fuels.
- But several factors delayed the development of solar technologies:
  - The Sun shines for only a part of each day.
  - In some areas clouds often reduce the amount of available radiation
- Different methods used to store the energy of the sun like **photovoltaic cells** like those used to supply power for astronauts in space.
- But these cells are not used for ordinary energy needs,because the cost is high comparing to the cost of burning of coal or oil.



SECTION 1 REVIEW

7. Chemical potential energy changes to heat in exothermic reactions, and the heat is released. In endothermic reactions, heat is absorbed and changed to chemical potential energy.
8. Two separated magnets illustrate potential energy. In a snow avalanche, positional potential energy changes to kinetic energy. Books on a shelf illustrate positional potential energy. As water races down a mountain stream, positional potential energy changes to kinetic energy. In a stock car race, chemical potential energy changes to kinetic energy. The separation of charge in a battery illustrates electrical potential energy.
9. Chemical potential energy, contained in the candle, is changed to energy in the form of light and heat and released as the chemical combustion reaction takes place.
10. 345 J
11. The temperature change is inversely proportional to the specific heat: aluminum, iron, silver, gold.

Solve the following questions :

1- Compare and contrast temperature and heat.

……………………………………………………………………………………………

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H.W: solve questions 53 – 67 .

In the space at the left, write *true* if the statement is true; if the statement is false, change the italicized word or phrase to make it true.

_______________ **1.** *Energy* is the ability to do work or produce heat.

_______________ **2.** The law of conservation of energy states that energy *can be* created and destroyed.

_______________ **3.** Chemical potential energy is energy stored in a substance because of its *composition*.

_______________ **4.** *Heat* is a form of energy that flows from a warmer object to a cooler object.

_______________ **5.** A calorie is the amount of energy required to raise the temperature of *one gram* of pure water by one degree Celsius.

_______________ **6.** A *calorie* is the SI unit of heat and energy.

_______________ **7.** The *specific heat* of a substance is the amount of heat required to raise the temperature of one gram of that substance by one degree Celsius.

_______________ **8.** *Kinetic energy* is energy of motion.

_______________ **9.** Chemicals participating in a chemical reaction contain *only potential energy*.

_______________ **10.** One nutritional Calorie is equal to *100 calories*.

_______________ **11.** One calorie equals *4.184 joules*.

_______________ **12.** When a fuel is burned, some of its *chemical potential energy* is lost as heat.

_______________ **13.** To convert kilojoules to joules, *divide* the number of kilojoules by 1000 joules/1 kilojoule.

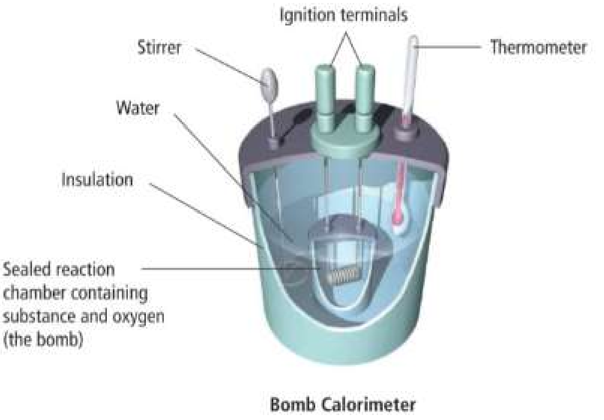
Section 2: Heat:

Calorimetry:

- Calorimeter: it is an **insulated** device used for measuring the amount of heat absorbed or released during a chemical or physical process.
- A bomb calorimeter : used by food chemists

Figure 5 A sample is positioned in a steel inner chamber called the bomb, which is filled with oxygen at high pressure. Surrounding the bomb is a measured mass of water stirred by a low-friction stirrer to ensure uniform temperature. The reaction is initiated by a spark, and the temperature is recorded until it reaches its maximum.

Infer *Why is it important that the stirrer does not create friction?*



Determining specific heat:

- In experiments we use simple foam cup calorimeter to determine the specific heat of unknown metal or any material.
  - Why is a foam cup used in a student calorimeter rather than a typical glass beaker?
    ……………………………………………………………………………………..
- These calorimeters are open to the atmosphere, so reactions carried out in them occure at constant pressure.

- <u>The experiment:</u>

Suppose you put 125 g of water into a foam-cup calorimeter and find that its initial temperature is 25.60°C. Then you heat a 50.0-g sample of the unknown metal to 115.0°C and put the metal sample into the water. Heat flows from the hot metal to the cooler water, and the temperature of the water rises. The flow of heat stops only when the temperature of the metal and the water are equal.

- The calculations: $q = c \times m \times \Delta T$
  - Calculate the heat gained by water:

$$q_{water} = 4.184\ J/(g\cdot°C) \times 125\ g \times (29.30°C - 25.60°C)$$
$$q_{water} = 4.184\ J/(g\cdot\cancel{°C}) \times 125\ \cancel{g} \times 3.70\cancel{°C}$$
$$q_{water} = 1940\ J$$

  - The heat gained by water = the heat lost by metal:

$$q_{metal} = q_{water}$$
$$q_{metal} = -1940\ J$$
$$c_{metal} \times m \times \Delta T = -1940\ J$$

  - Then solve the equation for specific heat of the metal:

$$c_{metal} = \frac{-1940\ J}{(50.0\ g)(-85.7°C)} = 0.453\ J/(g\cdot°C)$$

  - **<u>Table</u>** 2 shows that the metal could be iron.

$q = c \times m \times \Delta T$

PRACTICE Problems

12. A 90.0-g sample of an unknown metal absorbed 25.6 J of heat as its temperature increased 1.18°C. What is the specific heat of the metal?

13. The temperature of a sample of water increases from 20.0°C to 46.6°C as it absorbs 5650 J of heat. What is the mass of the sample?

14. How much heat is absorbed by a 2.00×10^3 g granite boulder ($c_{granite} = 0.803$ J/(g·°C)) as its temperature changes from 10.0°C to 29.0°C?

15. **Challenge** If 335 g of water at 65.5°C loses 9750 J of heat, what is the final temperature of the water?

0.241 J/(g·°C) .12
50.8 g .13
30,500 J .14
58.5°C .15

Question 75: A swimming pool measuring 20.0 m × 12.5 m is filled with water to a depth of 3.75 m. If the initial temperature is 18.4°C, how much heat must be added to the water to raise its temperature to 29.0°C? Assume that the density of water is 1.000 g/mL.

Chemical Energy and the Universe:

- **Thermochemistry**: is the study of heat changes that accompany chemical reactions and phase changes.
- Soldiers in the field use a highly exothermic reaction to heat their meals.
- You might have used a **heat back** to warm your hands on a cold day.

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s) + 1625 \text{ kJ}$$

- The **system**: is the specific part of the universe that contains the reaction or process you wish to study.
- The **surroundings**:everything in the universe other than the system.
- The **universe**:is the system plus the surrounding. $\text{universe} = \text{system} + \text{surroundings}$
- In the exothermic reaction heat flows from the system to the surrounding.
- In endothermic reaction heat flows from the surrounding to the system.

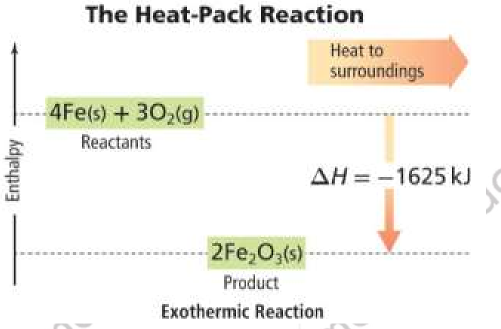
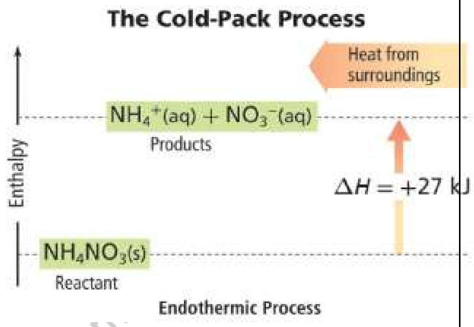
Figure 7 In this endothermic reaction, the reacting mixture draws enough energy from the water and the wet board to cause the wet board to freeze to the beaker.

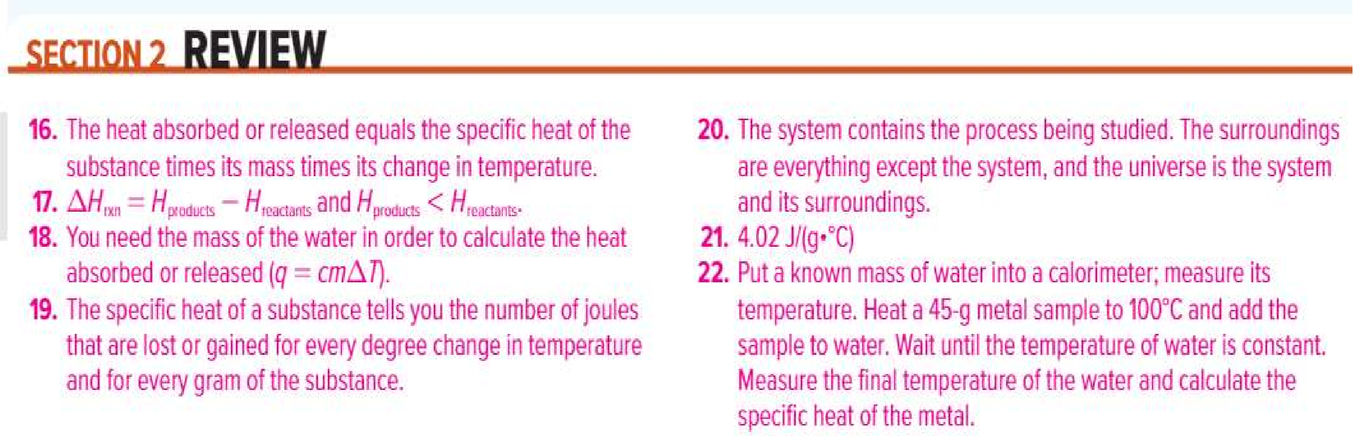
Enthalpy and enthalpy changes:

- The total amount of energy a substance contains depends on many factors , some of which are still not completely understood.
- Therefore , ti is impossible to know the total energy content of a substance.
- Therefore, chemists are usually more interested in changes in energy during reactions than in the absolute amounts of energy contained in the reactants and products.
- For many reactions , the amount of energy lost or gained can be measured in a calorimeter at constant pressure.
- The foam cup is not sealed so **the pressure is constant**.
- Many reactions take place at constant atmospheric pressure, for example, those that occur in living organisms.
- The energy released or evolved from reactions carried out at constant pressure is sometimes given the symbol (q_p).
- To more easily measure and study the energy changes, chemists have defined a property called **enthalpy**.
- **Enthalpy (H):** is the heat content of a system at constant pressure.
- **Enthalpy (heat) of reaction (ΔH):** the change in enthalpy for areaction.

$$\Delta H_{rxn} = H_{final} - H_{initial} \qquad \Delta H_{rxn} = H_{products} - H_{reactants}$$

- **Compare between exothermic and endothermic reaction:**

	Exothermic reaction **(hot pack)**	**Endothermic reaction** **(cold pack)**
Hreactants		
Hproducts		
The sign of ΔH		
Example	$4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)} + 1625\ kJ$ $4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}\ \Delta H_{rxn} = -1625\ kJ$	$27\ kJ + NH_4NO_{3(s)} \rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$ $NH_4NO_{3(s)} \rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}\ \ \Delta H_{rxn} = 27\ kJ$
Adiagram of the enthalpy change		



SECTION 2 REVIEW

16. The heat absorbed or released equals the specific heat of the substance times its mass times its change in temperature.
17. $\Delta H_{rxn} = H_{products} - H_{reactants}$ and $H_{products} < H_{reactants}$.
18. You need the mass of the water in order to calculate the heat absorbed or released ($q = cm\Delta T$).
19. The specific heat of a substance tells you the number of joules that are lost or gained for every degree change in temperature and for every gram of the substance.
20. The system contains the process being studied. The surroundings are everything except the system, and the universe is the system and its surroundings.
21. 4.02 J/(g•°C)
22. Put a known mass of water into a calorimeter; measure its temperature. Heat a 45-g metal sample to 100°C and add the sample to water. Wait until the temperature of water is constant. Measure the final temperature of the water and calculate the specific heat of the metal.

H.W:

- Solve the questions 68 – 78 .
- <u>**Describe a procedure**</u> you could follow to determine the specific heat of a 45-g piece of metal.

For each item in Column A, write the letter of the matching item in Column B.

Column A

_______ **1.** An insulated device used to measure the amount of heat absorbed or released during a chemical or physical process

_______ **2.** The study of heat changes that accompany chemical reactions and phase changes

_______ **3.** The specific part of the universe that contains the reaction or process you wish to study

_______ **4.** The change in enthalpy in a chemical reaction

_______ **5.** A system plus its surroundings

_______ **6.** The heat content of a system at constant pressure

_______ **7.** Everything in the universe except the system being studied

Column B

a. system

b. calorimeter

c. thermochemistry

d. universe

e. enthalpy

f. enthalpy (heat) of reaction

g. surroundings

Use the illustration to answer the following questions.

8. A scientist is studying the solution in the flask. What is the system?

____

9. What are the surroundings?

____

10. What is the universe?

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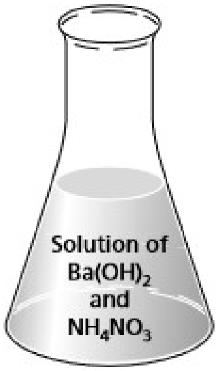

Section 3: Thermochemical Equations

Writing Thermochemical Equations:

- **A thermochemical equations**: is abalanced chemical equation that includes the physical states of all reactants and products and the energy changes ΔH.
- Examples :

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s) \quad \Delta H = -1625 \text{ kJ}$$

$$NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq) \quad \Delta H = 27 \text{ kJ}$$

- Combusition of glucose occurs in the body as food is metabolized to produce energy

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) \quad \Delta H_{comb} = -2808 \text{ kJ}$$

- The heat of combustion (ΔH_{comb}) of asubstance:is the enthalby change for the complete burning of 1 mol of the substance.
- Standard enthalpy changes have the symbol ΔH° (It is determined at standard conditions : 1 atm pressure and 25C°)

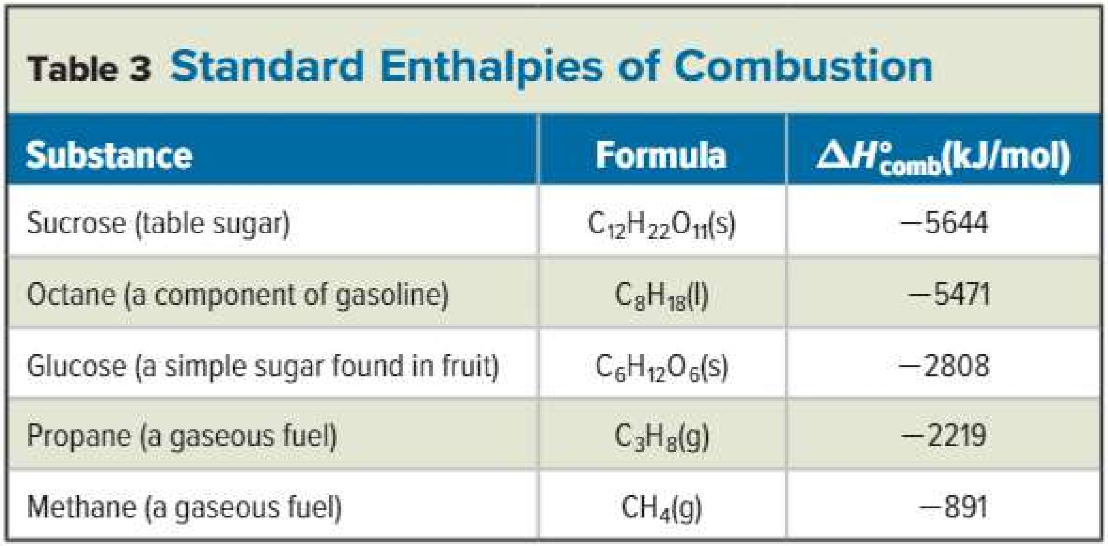

Table 3 Standard Enthalpies of Combustion

Substance	Formula	ΔH°_{comb}(kJ/mol)
Sucrose (table sugar)	$C_{12}H_{22}O_{11}(s)$	−5644
Octane (a component of gasoline)	$C_8H_{18}(l)$	−5471
Glucose (a simple sugar found in fruit)	$C_6H_{12}O_6(s)$	−2808
Propane (a gaseous fuel)	$C_3H_8(g)$	−2219
Methane (a gaseous fuel)	$CH_4(g)$	−891

Changes of states:

- When you step out of a hot shower you shiver as water evaporates from your skin? Because your skin provides the heat needed to vaporize the water.
- **Molar heat of vaporizatio**(ΔH_{vap}): the heat required to vaporize one mole of a liquid.
- **Molar heat of fusion** (ΔH_{fus}): the heat required to melt one mole of a solid substance (both of them are endothermic processes)

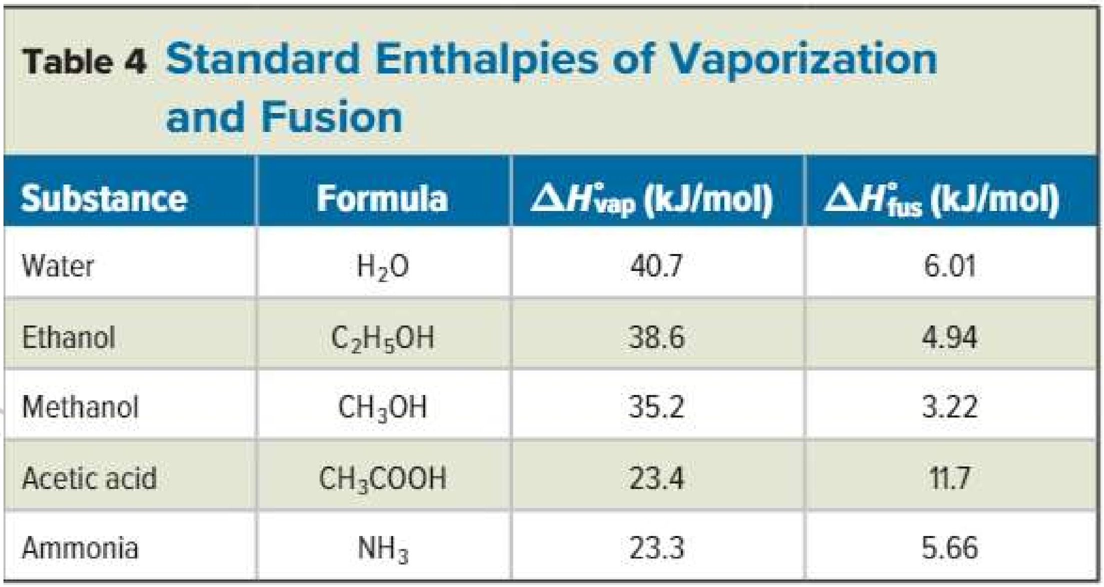

Table 4 Standard Enthalpies of Vaporization and Fusion

Substance	Formula	ΔH°_{vap} (kJ/mol)	ΔH°_{fus} (kJ/mol)
Water	H_2O	40.7	6.01
Ethanol	C_2H_5OH	38.6	4.94
Methanol	CH_3OH	35.2	3.22
Acetic acid	CH_3COOH	23.4	11.7
Ammonia	NH_3	23.3	5.66

- Thermochemical equations for changes of state:

$$H_2O_{(l)} \rightarrow H_2O_{(g)} \qquad \Delta H_{vap} = 40.7\ kJ$$

$$H_2O_{(s)} \rightarrow H_2O_{(l)} \qquad \Delta H_{fus} = 6.01\ kJ$$

- Question: what happens in the reverse processes?

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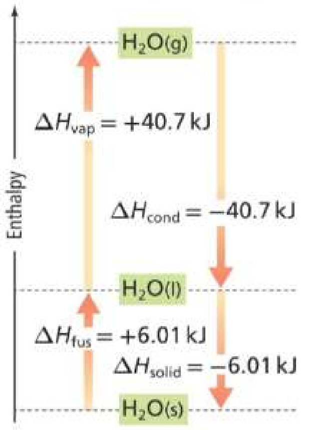
- So ΔH_{cond} and ΔH_{vap} have the same numerical value but opposite signs.
- ΔH_{solid} and ΔH_{fus} have the same numerical value but opposite signs.

$$\Delta H_{vap} = -\Delta H_{cond}$$

$$\Delta H_{fus} = -\Delta H_{solid}$$

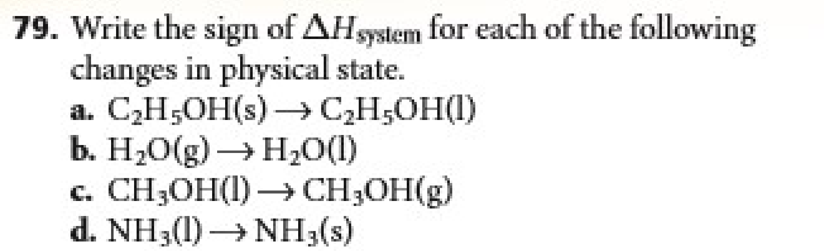
Figure 10 The upward arrows show that the energy of the system increases as ice melts and then vaporizes. The downward arrows show that the energy of the system decreases as water vapor condenses and then solidifies.

Phase Changes for Water



- Some farmers make use of the heat of fusion of water to protect fruit and vegetables from freezing , so they flood their fields with water. When the water freezes energy released ΔH_{solid} and often warms the surrounding air enough to prevent frost damage

Answer the following question :



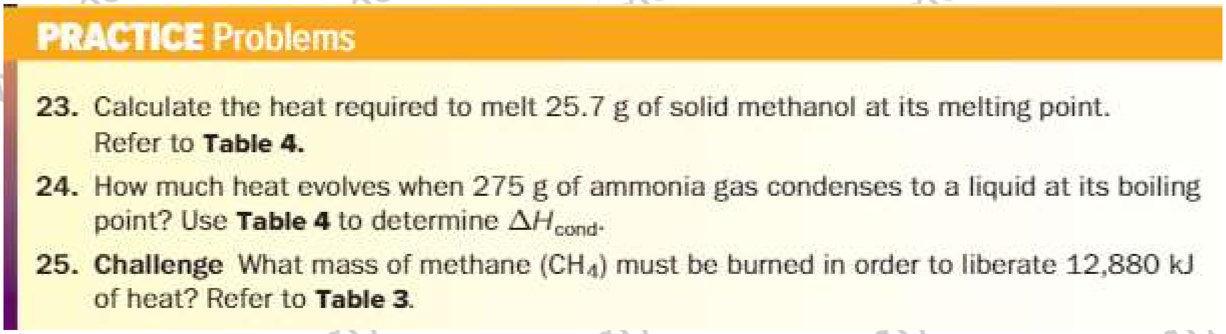
79. Write the sign of ΔH_{system} for each of the following changes in physical state.

a. $C_2H_5OH(s) \rightarrow C_2H_5OH(l)$

b. $H_2O(g) \rightarrow H_2O(l)$

c. $CH_3OH(l) \rightarrow CH_3OH(g)$

d. $NH_3(l) \rightarrow NH_3(s)$

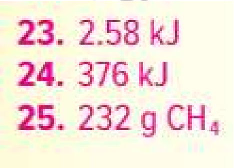


PRACTICE Problems

23. Calculate the heat required to melt 25.7 g of solid methanol at its melting point. Refer to **Table 4.**

24. How much heat evolves when 275 g of ammonia gas condenses to a liquid at its boiling point? Use **Table 4** to determine ΔH_{cond}.

25. **Challenge** What mass of methane (CH_4) must be burned in order to liberate 12,880 kJ of heat? Refer to **Table 3.**



23. 2.58 kJ

24. 376 kJ

25. 232 g CH_4

Combustion Reaction:

- Combustion : is the reaction of a fuel with oxygen.
- In biological systems food is the fuel (carbohydrates) which converted to glugose in your body.
- Other examples on combustion reactions:

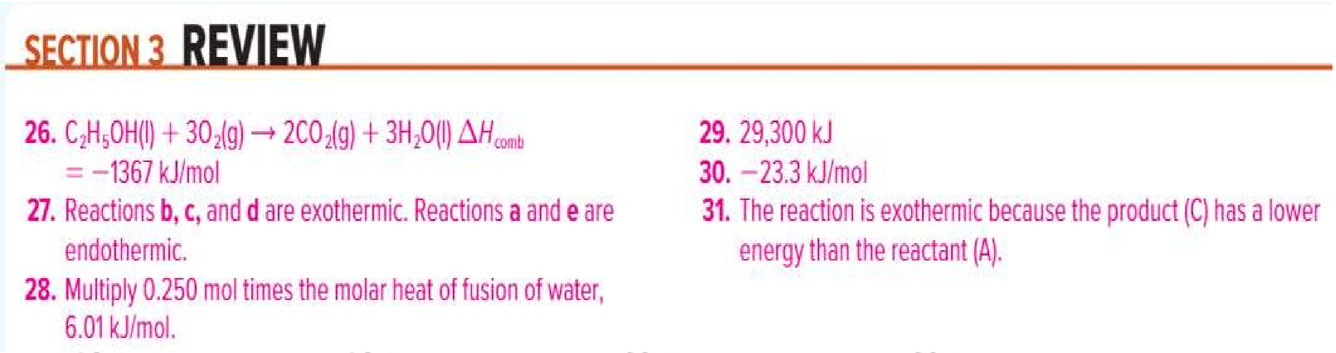
$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) + 891\ kJ$$

$$C_8H_{18(l)} + \frac{25}{2}O_{2(g)} \rightarrow 8CO_{2(g)} + 9H_2O_{(l)} + 5471\ kJ$$

- The combustion of hydrogen provide the energy to lift the shuttle into space:

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) + 286\ kJ$$

- **H.W:** Solve questions 79 – 88 .



SECTION 3 REVIEW

26. $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)\ \Delta H_{comb} = -1367$ kJ/mol

27. Reactions **b, c,** and **d** are exothermic. Reactions **a** and **e** are endothermic.

28. Multiply 0.250 mol times the molar heat of fusion of water, 6.01 kJ/mol.

29. 29,300 kJ

30. −23.3 kJ/mol

31. The reaction is exothermic because the product (C) has a lower energy than the reactant (A).

Use the following terms to complete the statements. Some terms will be used more than once.

- thermochemical equation
- molar enthalpy of vaporization
- cool
- enthalpy of combustion
- molar enthalpy of fusion
- heat
- released
- absorbs

1. A(n) ______________________________ is a balanced chemical equation that includes the physical states of all reactants and products and the energy change that accompanies the reaction.

2. The enthalpy change for the complete burning of one mole of a substance is the

______________________________.

3. The ______________________________ is the heat required to vaporize one mole of a liquid.

4. The ______________________________ is the heat required to melt one mole of a solid substance.

5. Converting two moles of a liquid to a solid requires an amount of energy that is twice

the ______________________________.

6. $2H_2(g) + O_2(g) \rightarrow 2H_2O(g) \quad \Delta H = -572$ kJ is a(n) ______________________________.

7. The conversion of a gas to a liquid involves the ______________________________.

8. When a gas condenses to a liquid, heat is ______________________________ to the surroundings.

9. Sweating makes you feel cooler because, as it evaporates, the water on your skin

______________________________ heat from your body.

10. If you put an ice cube in a glass of soda pop, the heat absorbed by the ice will cause the

ice to melt, and the soda pop will become ______________________________.

11. If it takes 100 joules to melt a piece of ice, ______________________________ must be absorbed by the ice.

12. In the equation $H_2O(s) \rightarrow H_2O(l) \quad \Delta H = 600$ kJ, the positive value for ΔH means that

______________________________ is absorbed in the reaction.

Section 4: Calculating Enthalpy Change

Hess's Law:

- **Sometimes it is impossible or impractical to measure the ΔH of a reaction by using a Calorimeter .**
  - **Some reactions occur so slowly that measuring the enthalpy change is impossible like the reaction of conversion of carbon in its allotropic form (diamond) to carbon in its allotropic form (graphite):** $C(s, \text{diamond}) \rightarrow C(s, \text{graphite})$
  - Other reactions occur under conditions difficult to duplicate in a laboratory .
  - Still others produce products other than the desired ones.
- For these reactions chemists use a theoretical way to determine ΔH .
- **Hess's Law:** states that if you can add tow or more thermochemical eqations to produce a final equation, then the sum of the enthalpy changes for the individual reactions is the enthalpy change for the final reaction.
- **Example:** use Hess's law to calculate the ΔH for the following reaction:

$$2S_{(s)} + 3O_{2(g)} \rightarrow 2SO_{3(g)} \quad \Delta H = ?$$

Using the following equations:

$$a.\ S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)} \quad \Delta H = -297\ kJ$$
$$b.\ 2SO_{3(g)} \rightarrow 2SO_{2(g)} + O_{2(g)} \quad \Delta H = 198\ kJ$$

Step1: multiply the first equation by 2:

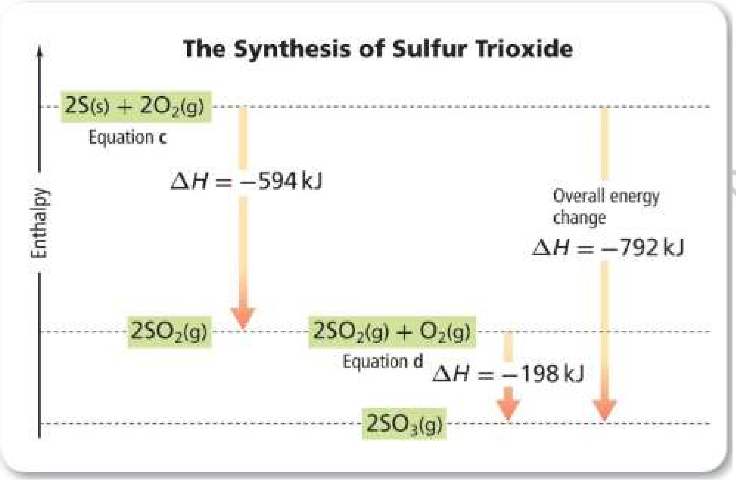
$$c.\ 2S_{(s)} + 2O_{2(g)} \rightarrow 2SO_{2(g)} \quad \Delta H = 2(-297\ kJ) = -594\ kJ$$

Step2: reverse the second equation:

$$d.\ 2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)} \quad \Delta H = -198\ kJ$$

Step3: add equations(c) and(d):

$$2S_{(s)} + 3O_{2(g)} \rightarrow 2SO_{3(g)} \quad \Delta H = -792\ kJ$$



HESS'S LAW Use thermochemical Equations **a** and **b** below to determine ΔH for the decomposition of hydrogen peroxide (H_2O_2), a compound that has many uses ranging from bleaching hair to powering rocket engines.

$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$

a. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)\ \Delta H = -572\ kJ$

b. $H_2(g) + O_2(g) \rightarrow H_2O_2(l)\ \Delta H = -188\ kJ$

$\Delta H = -196\ kJ$

Question Determine ΔH for the reaction C(s, graphite) + 1/2 $O_2(g) \rightarrow CO(g)$ using the following thermochemical equations.

A. C(s, graphite) + $O_2(g) \rightarrow CO_2(g)$
$\Delta H = -394$ kJ

B. $CO(g) + 1/2\ O_2(g) \rightarrow CO_2(g)$
$\Delta H = -283$ kJ

$\Delta H = -111\ kJ$

PRACTICE Problems

32. Use Equations **a** and **b** to determine ΔH for the following reaction.

$2CO(g) + 2NO(g) \rightarrow 2CO_2(g) + N_2(g)\ \ \Delta H = ?$

a. $2CO(g) + O_2(g) \rightarrow 2CO_2(g)\ \ \Delta H = -566.0\ kJ$

b. $N_2(g) + O_2(g) \rightarrow 2NO(g)\ \ \Delta H = -180.6\ kJ$

33. Challenge ΔH for the following reaction is -1789 kJ. Use this and Equation **a** to determine ΔH for Equation **b**.

$4Al(s) + 3MnO_2(s) \rightarrow 2Al_2O_3(s) + 3Mn(s)\ \ \Delta H = -1789\ kJ$

a. $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)\ \ \Delta H = -3352\ kJ$

b. $Mn(s) + O_2(g) \rightarrow MnO_2(s)\ \ \Delta H = ?$

32. -385.4 kJ

33. Because the direction of Equatio was changed, ΔH for Equation b $= -521$ kJ.

Standared heat of formation:

- Formation reaction: A reaction in which a compound is formed from its elements in their standared states
- The standared state of a substance means : the normal physical state of the substance at 1 atm and 298K (25C°) for example , iron is a solid , mercury is a liquid , oxygen is a diatomic gas.
- **The stadared heat of formation(ΔH_f°)**: the change in enthalpy that accompanies the formation of one mole of the compound in its standared state from its elements in their standared states.
- **For example**: formation of one mol of SO_3 :
- SO_3 is a suffocating gas , produces acid rain when mix with moisture in the atmosphere .

$$S(s) + \frac{3}{2}O_2(g) \rightarrow SO_3(g) \quad \Delta H_f^\circ = -396 \text{ kJ}$$

Figure 14 Sulfur trioxide combines with water in the atmosphere to form sulfuric acid (H_2SO_4), a strong acid, which reaches Earth as acid precipitation. Acid precipitation slowly destroys trees and property.

Where do standared heat of formation come from?

- Elements in their standared state have a **ΔH_f° of 0.0 kJ**
- Standared heat of formation of many compounds have been measured experimentally:

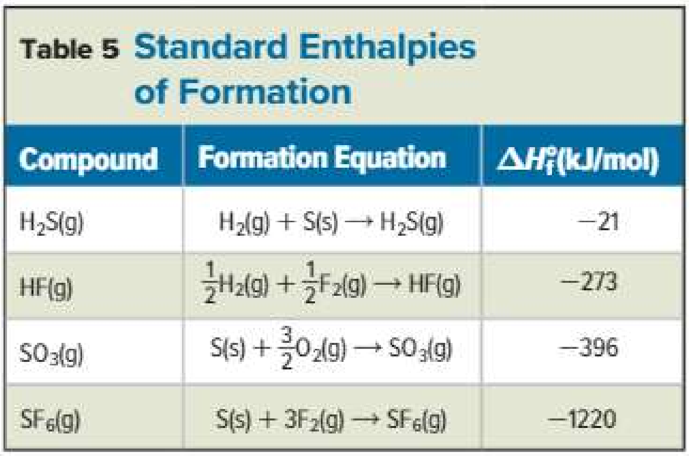


Table 5 Standard Enthalpies of Formation

Compound	Formation Equation	ΔH_f°(kJ/mol)
$H_2S(g)$	$H_2(g) + S(s) \rightarrow H_2S(g)$	−21
$HF(g)$	$\frac{1}{2}H_2(g) + \frac{1}{2}F_2(g) \rightarrow HF(g)$	−273
$SO_3(g)$	$S(s) + \frac{3}{2}O_2(g) \rightarrow SO_3(g)$	−396
$SF_6(g)$	$S(s) + 3F_2(g) \rightarrow SF_6(g)$	−1220

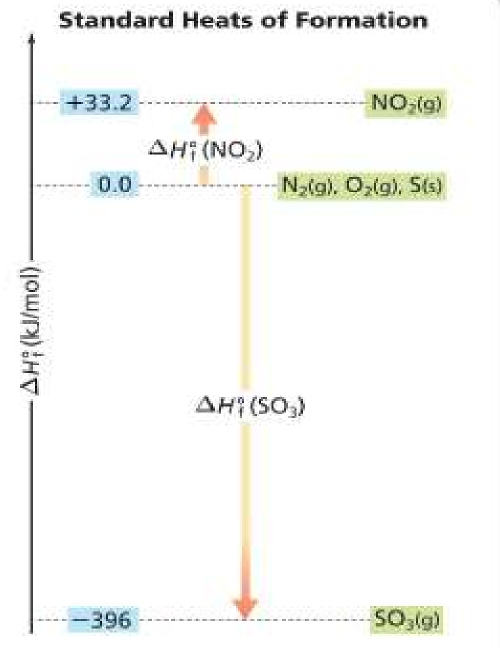


Figure 15 ΔH_f° for the elements N_2, O_2, and S is 0.0 kJ. When N_2 and O_2 react to form 1 mole of NO_2, 33.2 kJ is absorbed. Thus, ΔH_f° for NO_2 is +33.2 kJ/mol. When S and O_2 react to form one mole of SO_3, 396 kJ is released. Therefore, ΔH_f° for SO_3 is −396 kJ/mol.

Predict *Describe the approximate location of water on the scale. The heat of formation for the reaction* $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ *is* $\Delta H_f^\circ = -286$ *kJ/mol.*

Using standared enthalpies (heat) of formation:

- It is used to calculate the enthalpies of many reactions using Hess's Law.
- Using table 5 , calculate ΔH°_{rxn} for the following reaction:

$$H_2S_{(g)} + 4F_{2(g)} \rightarrow 2HF_{(g)} + SF_{6(g)} \qquad \Delta H^\circ_{rxn} = ?$$

- Refer to table 5 to find an equation of the formation of each of the three compounds:

a. $\frac{1}{2}H_2(g) + \frac{1}{2}F_2(g) \rightarrow HF(g) \qquad \Delta H^\circ_f = -273\ kJ$
b. $S(s) + 3F_2(g) \rightarrow SF_6(g) \qquad \Delta H^\circ_f = -1220\ kJ$
c. $H_2(g) + S(s) \rightarrow H_2S(g) \qquad \Delta H^\circ_f = -21\ kJ$

Figure 16 Sulfur hexafluoride is used to etch minute and sometimes intricate patterns on silicon wafers in the production semiconductor devices. Semiconductors are important components of modern electronic equipment, including computers, cell phones, and MP3 players.

$$\boxed{\Delta H^\circ_{rxn} = -1745\ kJ}$$

Complete the solution:

----

94. Use Hess's law and the changes in enthalpy for the following two generic reactions to calculate ΔH for the reaction $2A + B_2C_3 \rightarrow 2B + A_2C_3$.

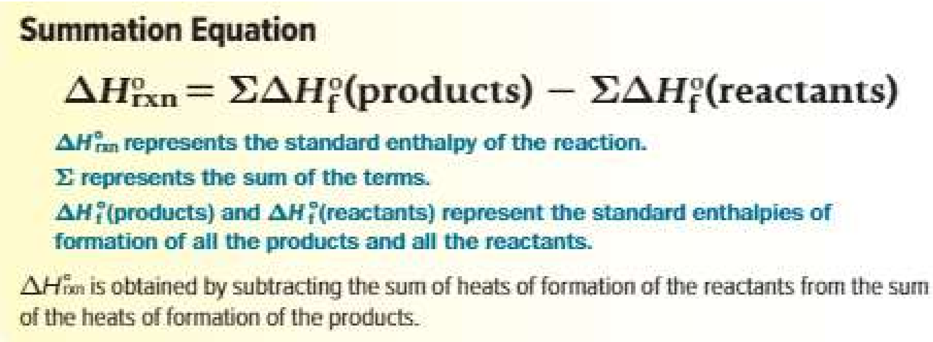
$$2A + \frac{3}{2}C_2 \rightarrow A_2C_3 \quad \Delta H = -1874\ kJ$$

$$2B + \frac{3}{2}C_2 \rightarrow B_2C_3 \quad \Delta H = -285\ kJ$$

109. Apply Phosphorus trichloride is a starting material for the preparation of organic phosphorous compounds. Demonstrate how thermochemical equations a and b can be used to determine the enthalpy change for the following reaction $PCl_3(l) + Cl_2(g) \rightarrow PCl_5(s)$.

a. $P_4(s) + 6Cl_2(g) \rightarrow 4PCl_3(l) \qquad \Delta H = -1280\ kJ$
b. $P_4(s) + 10Cl_2(g) \rightarrow 4PCl_5(s) \qquad \Delta H = -1774\ kJ$

The summation equation:



Summation Equation

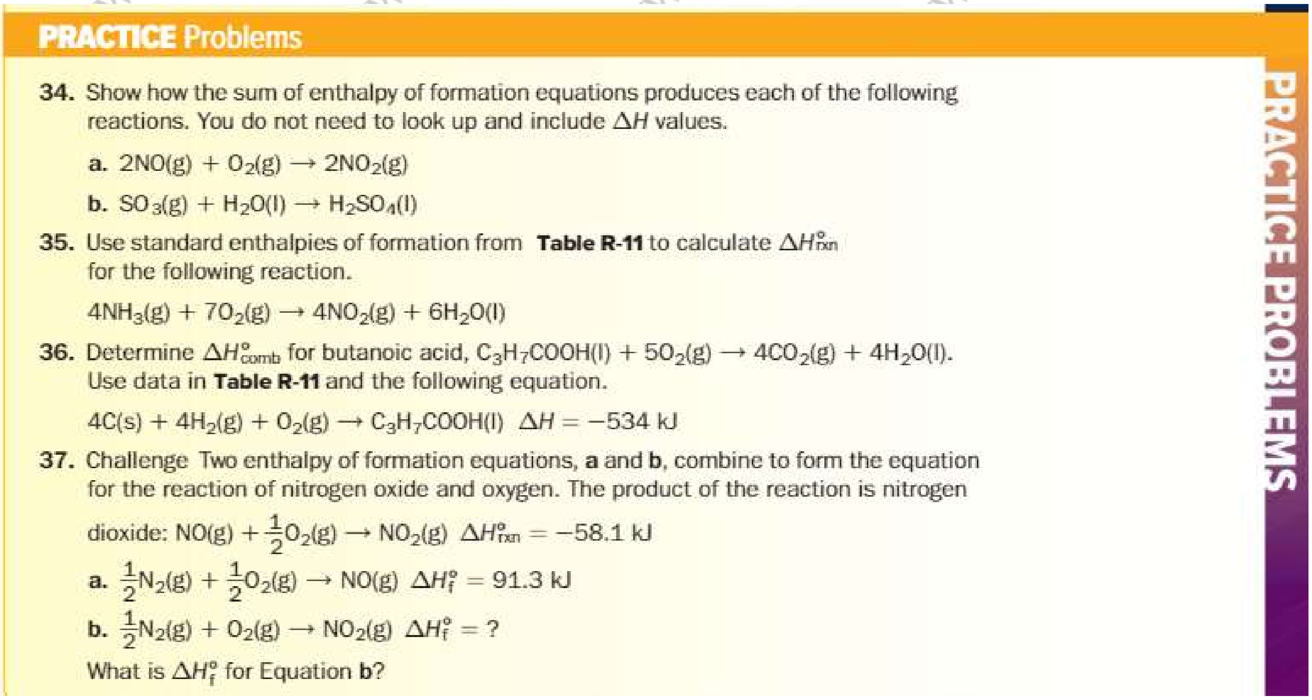
$$\Delta H^\circ_{rxn} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$$

ΔH°_{rxn} represents the standard enthalpy of the reaction.

Σ represents the sum of the terms.

ΔH°_f(products) and ΔH°_f(reactants) represent the standard enthalpies of formation of all the products and all the reactants.

ΔH°_{rxn} is obtained by subtracting the sum of heats of formation of the reactants from the sum of the heats of formation of the products.



PRACTICE Problems

34. Show how the sum of enthalpy of formation equations produces each of the following reactions. You do not need to look up and include ΔH values.

a. $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

b. $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l)$

35. Use standard enthalpies of formation from **Table R-11** to calculate ΔH°_{rxn} for the following reaction.

$4NH_3(g) + 7O_2(g) \rightarrow 4NO_2(g) + 6H_2O(l)$

36. Determine ΔH°_{comb} for butanoic acid, $C_3H_7COOH(l) + 5O_2(g) \rightarrow 4CO_2(g) + 4H_2O(l)$. Use data in **Table R-11** and the following equation.

$4C(s) + 4H_2(g) + O_2(g) \rightarrow C_3H_7COOH(l) \quad \Delta H = -534\text{ kJ}$

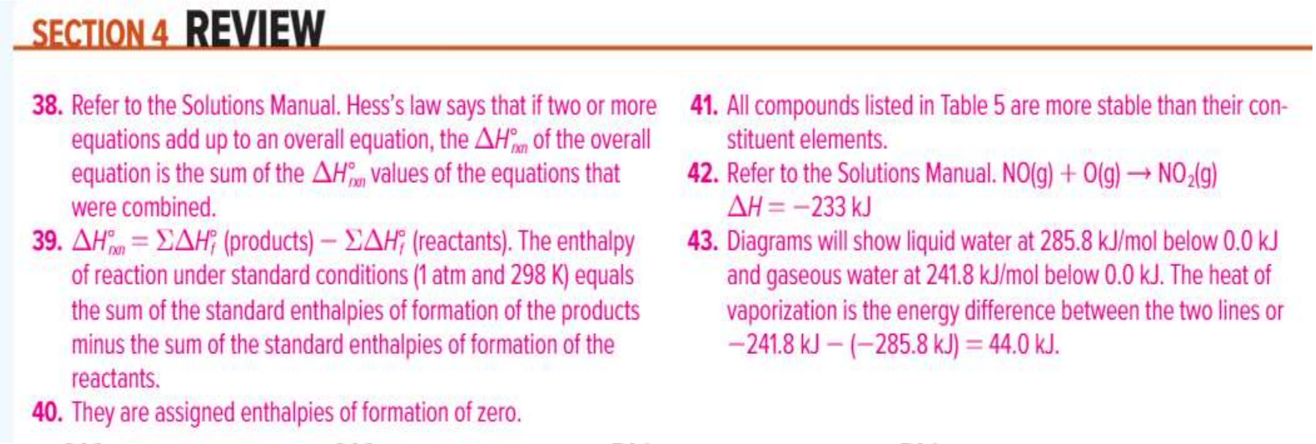
37. Challenge Two enthalpy of formation equations, **a** and **b**, combine to form the equation for the reaction of nitrogen oxide and oxygen. The product of the reaction is nitrogen dioxide: $NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g) \quad \Delta H^\circ_{rxn} = -58.1\text{ kJ}$

a. $\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightarrow NO(g) \quad \Delta H^\circ_f = 91.3\text{ kJ}$

b. $\frac{1}{2}N_2(g) + O_2(g) \rightarrow NO_2(g) \quad \Delta H^\circ_f = ?$

What is ΔH°_f for Equation **b**?

PRACTICE PROBLEMS



SECTION 4 REVIEW

38. Refer to the Solutions Manual. Hess's law says that if two or more equations add up to an overall equation, the ΔH°_{rxn} of the overall equation is the sum of the ΔH°_{rxn} values of the equations that were combined.

39. $\Delta H^\circ_{rxn} = \Sigma\Delta H^\circ_f$ (products) − $\Sigma\Delta H^\circ_f$ (reactants). The enthalpy of reaction under standard conditions (1 atm and 298 K) equals the sum of the standard enthalpies of formation of the products minus the sum of the standard enthalpies of formation of the reactants.

40. They are assigned enthalpies of formation of zero.

41. All compounds listed in Table 5 are more stable than their constituent elements.

42. Refer to the Solutions Manual. $NO(g) + O(g) \rightarrow NO_2(g)$ $\Delta H = -233\text{ kJ}$

43. Diagrams will show liquid water at 285.8 kJ/mol below 0.0 kJ and gaseous water at 241.8 kJ/mol below 0.0 kJ. The heat of vaporization is the energy difference between the two lines or $-241.8\text{ kJ} - (-285.8\text{ kJ}) = 44.0\text{ kJ}$.

H.W: solve questions 89 – 94 .

In the space at the left, write *true* if the statement is true; if the statement is false, change the italicized word or phrase to make it true.

____________________ **1.** Hess's law states that if two or more thermochemical equations can be added to produce a final equation for a reaction, then the sum of all the enthalpy changes for the individual reactions is the enthalpy change for the *final reaction*.

____________________ **2.** The standard enthalpy of formation is the change in enthalpy that accompanies the formation of *one gram* of a compound in its standard state from its constituent elements in their standard states.

____________________ **3.** The standard state of iron is *solid*.

____________________ **4.** For a pure gas, the standard state is the gas at a pressure of *one atmosphere*.

____________________ **5.** The symbol used to represent standard enthalpy of formation is ΔH_f°.

____________________ **6.** The standard state of a substance is the normal state of the substance at *0 K* and one atmosphere pressure.

____________________ **7.** The standard enthalpy of formation of a free element in its standard state is *0.0 kJ*.

____________________ **8.** A standard enthalpy of formation that has a *negative* value means that energy is absorbed during the reaction.

____________________ **9.** The standard state of oxygen is *gas*.

____________________ **10.** Standard enthalpies of formation provide data for calculating the enthalpies of reactions under standard conditions using *Hess's law*.

____________________ **11.** The standard state of mercury is *solid*.

Reference Tables:

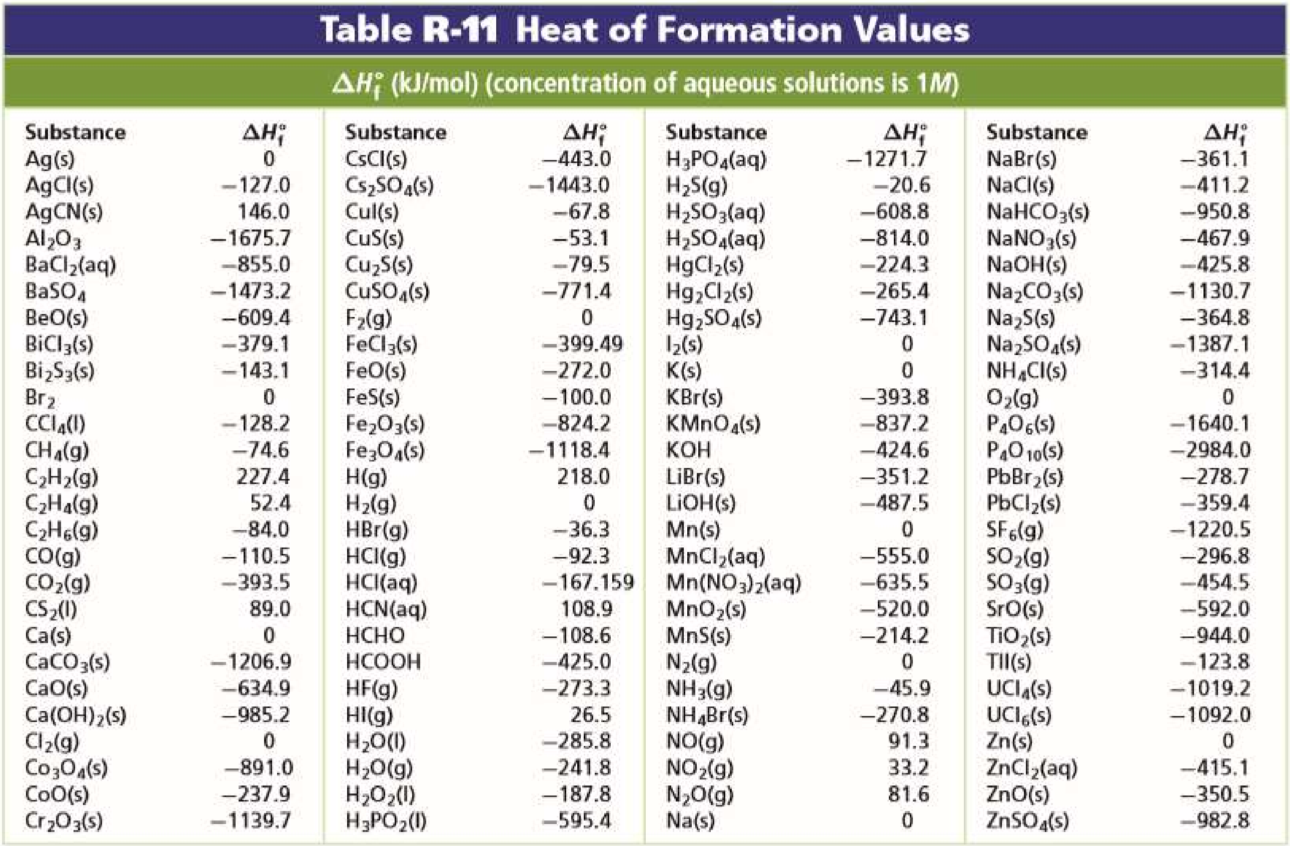


Table R-11 Heat of Formation Values

ΔH_f° (kJ/mol) (concentration of aqueous solutions is 1*M*)

Substance	ΔH_f°	Substance	ΔH_f°	Substance	ΔH_f°	Substance	ΔH_f°
$Ag(s)$	0	$CsCl(s)$	−443.0	$H_3PO_4(aq)$	−1271.7	$NaBr(s)$	−361.1
$AgCl(s)$	−127.0	$Cs_2SO_4(s)$	−1443.0	$H_2S(g)$	−20.6	$NaCl(s)$	−411.2
$AgCN(s)$	146.0	$CuI(s)$	−67.8	$H_2SO_3(aq)$	−608.8	$NaHCO_3(s)$	−950.8
Al_2O_3	−1675.7	$CuS(s)$	−53.1	$H_2SO_4(aq)$	−814.0	$NaNO_3(s)$	−467.9
$BaCl_2(aq)$	−855.0	$Cu_2S(s)$	−79.5	$HgCl_2(s)$	−224.3	$NaOH(s)$	−425.8
$BaSO_4$	−1473.2	$CuSO_4(s)$	−771.4	$Hg_2Cl_2(s)$	−265.4	$Na_2CO_3(s)$	−1130.7
$BeO(s)$	−609.4	$F_2(g)$	0	$Hg_2SO_4(s)$	−743.1	$Na_2S(s)$	−364.8
$BiCl_3(s)$	−379.1	$FeCl_3(s)$	−399.49	$I_2(s)$	0	$Na_2SO_4(s)$	−1387.1
$Bi_2S_3(s)$	−143.1	$FeO(s)$	−272.0	$K(s)$	0	$NH_4Cl(s)$	−314.4
Br_2	0	$FeS(s)$	−100.0	$KBr(s)$	−393.8	$O_2(g)$	0
$CCl_4(l)$	−128.2	$Fe_2O_3(s)$	−824.2	$KMnO_4(s)$	−837.2	$P_4O_6(s)$	−1640.1
$CH_4(g)$	−74.6	$Fe_3O_4(s)$	−1118.4	KOH	−424.6	$P_4O_{10}(s)$	−2984.0
$C_2H_2(g)$	227.4	$H(g)$	218.0	$LiBr(s)$	−351.2	$PbBr_2(s)$	−278.7
$C_2H_4(g)$	52.4	$H_2(g)$	0	$LiOH(s)$	−487.5	$PbCl_2(s)$	−359.4
$C_2H_6(g)$	−84.0	$HBr(g)$	−36.3	$Mn(s)$	0	$SF_6(g)$	−1220.5
$CO(g)$	−110.5	$HCl(g)$	−92.3	$MnCl_2(aq)$	−555.0	$SO_2(g)$	−296.8
$CO_2(g)$	−393.5	$HCl(aq)$	−167.159	$Mn(NO_3)_2(aq)$	−635.5	$SO_3(g)$	−454.5
$CS_2(l)$	89.0	$HCN(aq)$	108.9	$MnO_2(s)$	−520.0	$SrO(s)$	−592.0
$Ca(s)$	0	$HCHO$	−108.6	$MnS(s)$	−214.2	$TiO_2(s)$	−944.0
$CaCO_3(s)$	−1206.9	$HCOOH$	−425.0	$N_2(g)$	0	$TlI(s)$	−123.8
$CaO(s)$	−634.9	$HF(g)$	−273.3	$NH_3(g)$	−45.9	$UCl_4(s)$	−1019.2
$Ca(OH)_2(s)$	−985.2	$HI(g)$	26.5	$NH_4Br(s)$	−270.8	$UCl_6(s)$	−1092.0
$Cl_2(g)$	0	$H_2O(l)$	−285.8	$NO(g)$	91.3	$Zn(s)$	0
$Co_3O_4(s)$	−891.0	$H_2O(g)$	−241.8	$NO_2(g)$	33.2	$ZnCl_2(aq)$	−415.1
$CoO(s)$	−237.9	$H_2O_2(l)$	−187.8	$N_2O(g)$	81.6	$ZnO(s)$	−350.5
$Cr_2O_3(s)$	−1139.7	$H_3PO_2(l)$	−595.4	$Na(s)$	0	$ZnSO_4(s)$	−982.8

Section 5: reaction spontaneity:

Spontaneous Processes:

- It is any physical or chemical change that once occurs with no outside intervention.
- All exothermic processes are spontaneous like formation of iron rust:
  $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s) \quad \Delta H = -1625\ kJ$
- All endothermic processes are nonspontaneous , like the reverse reaction of iron dusting: $2Fe_2O_3(s) \rightarrow 4Fe(s) + 3O_2(g) \quad \Delta H = 1625\ kJ$
- Ice melting at room tempreture spontaneous , endothermic process, this mean that something other than ΔH plays a role in determining weather a chemical process occurs spontaneously like entropy.

What is entropy?

- Entropy (S) :Is a measure of the number of possible ways that the energy of a system can be distributed.
- This is related to the freedom of the system's particles to move and number of ways they can be arranged

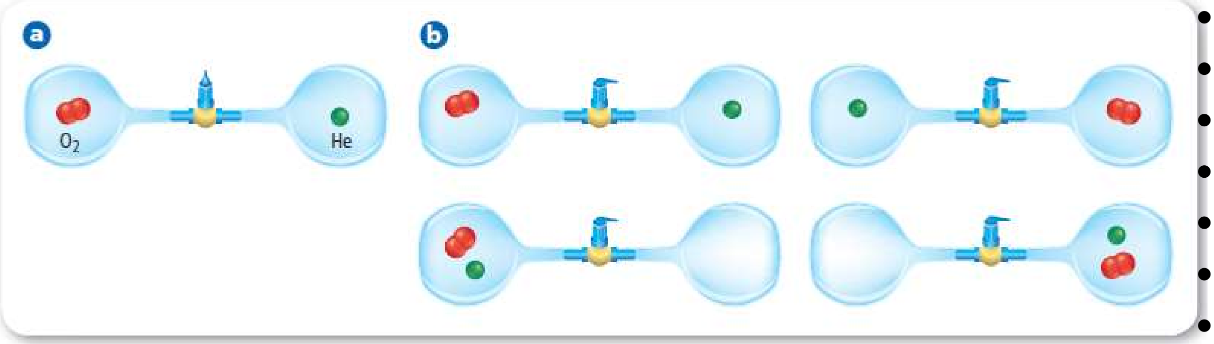


Figure 18 In **a,** an oxygen molecule and a helium atom are each confined to a single bulb. When the stopcock is opened in **b,** the gas particles move freely into the double volume available. Four arrangements of the particles, which represent an increase in entropy, are possible at any given time.

- As the number of particles increases , the number of possible arrangements for a group of particles increases.
- If the tow bulbs contained a total of **ten** particles , the number of possible arrangements would be 1024 times.
- In general, the number of possible arrangements available to a system increases under the following conditions:
  - When volume increases.
  - When energy increases.
  - When the number of particles increases.
  - When the particles' freedom of movement increases.

The second law of thermodynamics:

- Spontaneous processes always proceed in such a way that the entropy of the universe increases.
- **Entropy : can be considered to be a measure of the disorder or randomness of the particles that make up a system.**

Predicting changes in entropy:

$$\Delta S_{system} = S_{products} - S_{reactants}$$

- If the entropy of a system increases $S_{products} > S_{reactants}$ and ΔS_{system} is positive.
- To predict if ΔS_{system} is positive or negative :

1- Changes in state: the movement of molecules in gases > liquids > solids , so the entropy will increases as the movement of molecules increases:

$$H_2O(l) \longrightarrow H_2O(g) \quad \Delta S_{system} > 0$$

$$CH_3OH(s) \longrightarrow CH_3OH(l) \quad \Delta S_{system} > 0$$

2- The dissolving of a gas in a solvent always results in a decrease in intropy:

$$O_2(g) \longrightarrow O_2(aq) \quad \Delta S_{system} < 0$$

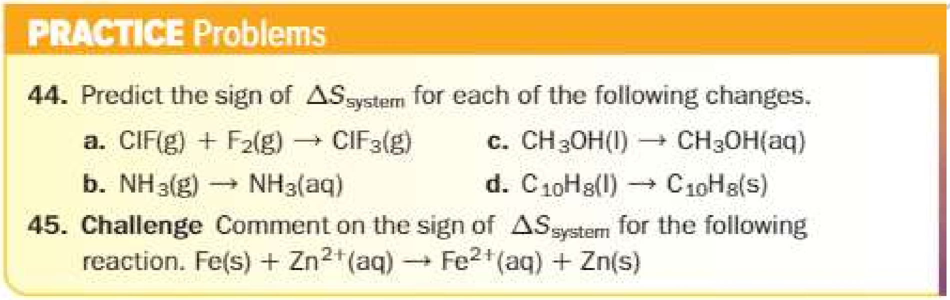
3- The intropy of a system usually increases when the number of gaseous products particles is greater than the number of gaseous reactant particles:

$$2SO_3(g) \rightarrow 2SO_2(g) + O_2(g) \quad \Delta S_{system} > 0$$

4- Intropy increases when a solid or a liquid dissolves in a solvent:

$$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq) \quad \Delta S_{system} > 0$$

5- The random motion of the particles of a substance increases as its temperature increases , therefore the intropy increases.



PRACTICE Problems

44. Predict the sign of ΔS_{system} for each of the following changes.

a. $ClF(g) + F_2(g) \rightarrow ClF_3(g)$
b. $NH_3(g) \rightarrow NH_3(aq)$
c. $CH_3OH(l) \rightarrow CH_3OH(aq)$
d. $C_{10}H_8(l) \rightarrow C_{10}H_8(s)$

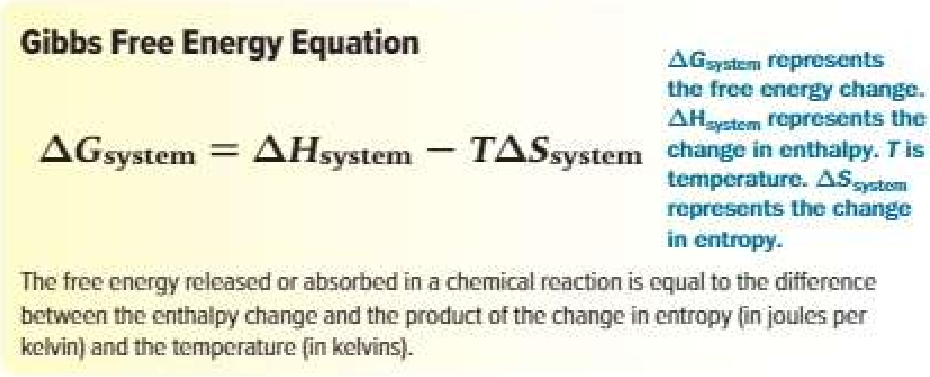
45. Challenge Comment on the sign of ΔS_{system} for the following reaction. $Fe(s) + Zn^{2+}(aq) \rightarrow Fe^{2+}(aq) + Zn(s)$

Entropy , the universe .and freedom: $\Delta S_{universe} = \Delta S_{system} + \Delta S_{surroundings}$

- In nature $\Delta S_{universe}$ tends to be positive for reactions and processes under the following conditions:
  1- The reaction is exothermic (ΔH_{system} is negative). This will raise the tempreture of the surroundings and increase the entropy of the surroundings (ΔS is positive).
  2- The entropy of the system increases . so ΔS_{system} is positive.
- Thus exothermic chemical reactions accompanied by an increase in entropy are all spontaneous.

Free Energy G :

- Gibbs free energy a combined enthalby entropy function.
- Free energy: is energy that is available to do work, thus free energy is useful energy, in contrast. Some entropy is associated with energy that is spread out into the surroundings, so it is useless.



Gibbs Free Energy Equation

$$\Delta G_{system} = \Delta H_{system} - T\Delta S_{system}$$

ΔG_{system} represents the free energy change. ΔH_{system} represents the change in enthalpy. T is temperature. ΔS_{system} represents the change in entropy.

The free energy released or absorbed in a chemical reaction is equal to the difference between the enthalpy change and the product of the change in entropy (in joules per kelvin) and the temperature (in kelvins).

- ΔS in J/K , ΔH IN kJ .
- If the sign of the free energy is negative , the reaction is spontaneous.
- If it is positive , the reaction is nonspontaneous.

Calculating free energy change:

- **Example:** calculate ΔG for the following system: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

$$\Delta H^\circ_{\text{system}} = -91.8\ \text{kJ} \quad \Delta S^\circ_{\text{system}} = -197\ \text{J/K}$$

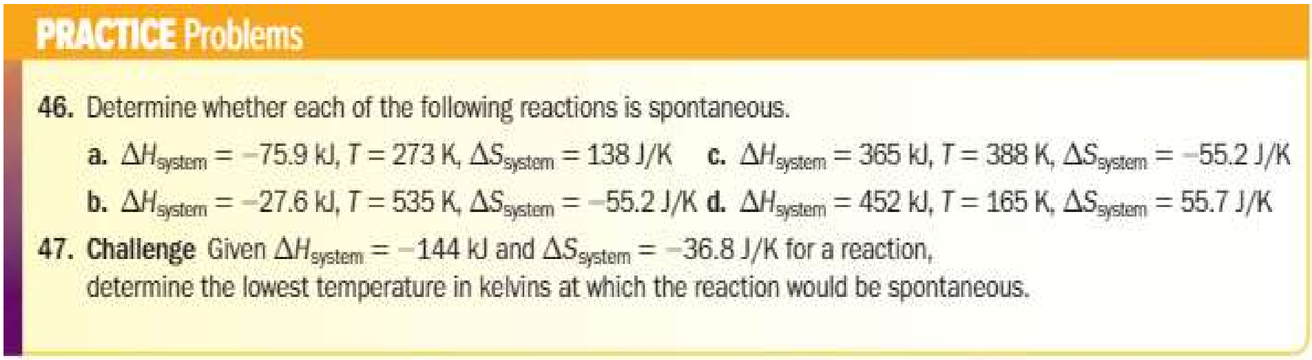
- Convert ΔS to kilojouls : $\Delta S^\circ_{\text{system}} = -197\ \text{J/K} \times \frac{1\ \text{kJ}}{1000\ \text{J}} = -0.197\ \text{kJ/K}$

-
$$\Delta G^\circ_{\text{system}} = \Delta H^\circ_{\text{system}} - T\Delta S^\circ_{\text{system}}$$
$$\Delta G^\circ_{\text{system}} = -91.8\ \text{kJ} - (298\ \cancel{\text{K}})(-0.197\ \text{kJ}/\cancel{\text{K}})$$
$$\Delta G^\circ_{\text{system}} = -91.8\ \text{kJ} + 58.7\ \text{kJ} = -33.1\ \text{kJ}$$

- ΔG is negative , so the reaction is spontaneous.
- The following table shows how reaction spontaneity depends on the sign of ΔH and ΔS :

Table 6 Reaction Spontaneity $\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$

ΔH_{system}	ΔS_{system}	ΔG_{system}	Reaction Spontaneity
negative	positive	always negative	always spontaneous
negative	negative	negative or positive	spontaneous at lower temperatures
positive	positive	negative or positive	spontaneous at higher temperatures
positive	negative	always positive	never spontaneous



PRACTICE Problems

46. Determine whether each of the following reactions is spontaneous.

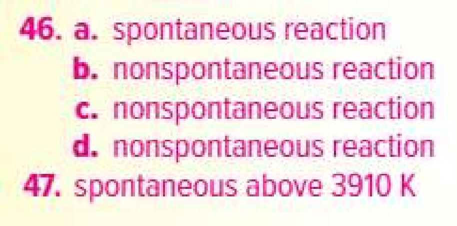
a. $\Delta H_{system} = -75.9$ kJ, $T = 273$ K, $\Delta S_{system} = 138$ J/K

b. $\Delta H_{system} = -27.6$ kJ, $T = 535$ K, $\Delta S_{system} = -55.2$ J/K

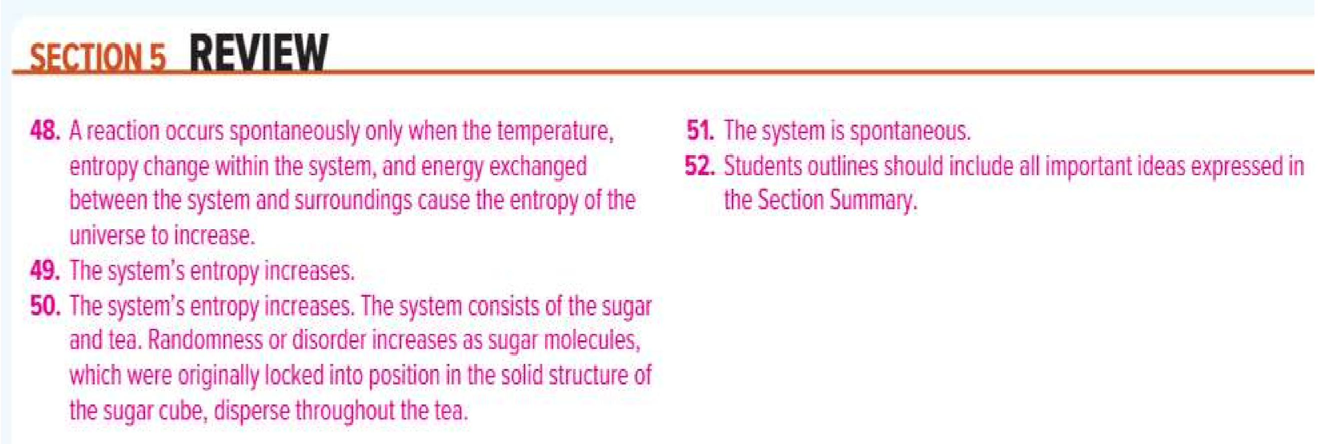
c. $\Delta H_{system} = 365$ kJ, $T = 388$ K, $\Delta S_{system} = -55.2$ J/K

d. $\Delta H_{system} = 452$ kJ, $T = 165$ K, $\Delta S_{system} = 55.7$ J/K

47. Challenge Given $\Delta H_{system} = -144$ kJ and $\Delta S_{system} = -36.8$ J/K for a reaction, determine the lowest temperature in kelvins at which the reaction would be spontaneous.



46. a. spontaneous reaction
b. nonspontaneous reaction
c. nonspontaneous reaction
d. nonspontaneous reaction
47. spontaneous above 3910 K



SECTION 5 REVIEW

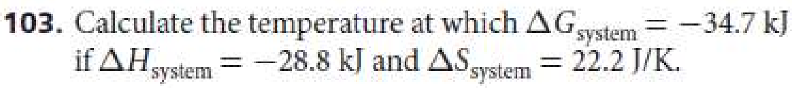
48. A reaction occurs spontaneously only when the temperature, entropy change within the system, and energy exchanged between the system and surroundings cause the entropy of the universe to increase.

49. The system's entropy increases.

50. The system's entropy increases. The system consists of the sugar and tea. Randomness or disorder increases as sugar molecules, which were originally locked into position in the solid structure of the sugar cube, disperse throughout the tea.

51. The system is spontaneous.

52. Students outlines should include all important ideas expressed in the Section Summary.



103. Calculate the temperature at which $\Delta G_{system} = -34.7$ kJ if $\Delta H_{system} = -28.8$ kJ and $\Delta S_{system} = 22.2$ J/K.

H.W : solve the questions 95 – 107

Use each of the terms below to complete the statements.

free energy	law of disorder	entropy	spontaneous process

1. A(n)______________________is a physical or chemical change that occurs with no outside intervention.

2. A measure of disorder or randomness of the particles that make up a system is called

______________________.

3. The______________________states that spontaneous processes always proceed in such a way that the entropy of the universe increases.

4. ______________________ is the energy that is available to do work.

For each statement below, write *true* or *false*.

_____________ **5.** A process cannot be spontaneous if it is exothermic and there is an increase in disorder.

_____________ **6.** A process cannot be spontaneous if it is endothermic and there is a decrease in disorder.

_____________ **7.** A process cannot be spontaneous if it is exothermic and there is a decrease in disorder as long as the temperature remains low.

_____________ **8.** A process cannot be spontaneous if it is endothermic and there is an increase in disorder as long as the temperature remains high.

_____________ **9.** A process can never be spontaneous if the entropy of the universe increases.

_____________ **10.** When A*G* for a reaction is negative, the reaction is spontaneous.

_____________ **11.** When A*G* for a reaction is positive, the reaction is not spontaneous.

_____________ **12.** When A*H* for a reaction is negative, the reaction is never spontaneous.

_____________ **13.** When A*H* for a reaction is large and positive, the reaction is not expected to be spontaneous.