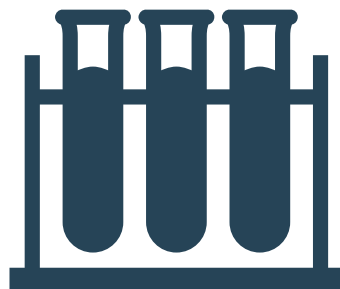


Chapter 1

Energy and Chemical Change



Revision Paper



مكتبة الفكر

Energy

Chapter 1 Lesson 1 – Revision Paper

Table 2

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	

Table 2 – Page 08

[Applications 4](#) - [Applications 5](#) – [Review 11](#)

1.1 Energy

Example 1

01

A breakfast of cereal, orange juice, and milk might contain **230 nutritional Calories**. Express this energy in joules.

Example 1

01

A breakfast of cereal, orange juice, and milk might contain **230 nutritional Calories**. Express this energy in joules.

Convert nutritional Calories to calories.

$$230 \text{ ~~Calories~~} \times \frac{1000 \text{ cal}}{1 \text{ ~~Calorie~~}} = 2.3 \times 10^5 \text{ cal}$$

Apply the relationship 1 Calorie = 1000 cal.

Convert calories to joules.

$$2.3 \times 10^5 \text{ ~~cal~~} \times \frac{4.184 \text{ J}}{1 \text{ ~~cal~~}} = 9.6 \times 10^5 \text{ J}$$

Apply the relationship 1 cal = 4.184 J.

Applications 1

02

A fruit-and-oatmeal bar contains **142 nutritional Calories**. Convert this energy to calories.

Applications 1

02

A fruit-and-oatmeal bar contains **142 nutritional Calories**. Convert this energy to calories.

$$142 \text{ Calories} = 142 \text{ kcal}$$

$$142 \text{ kcal} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} = 142,000 \text{ cal}$$

Applications 2

03

An exothermic reaction releases **86.5 kJ**. How many kilocalories of energy are released?

Applications 2

03

An exothermic reaction releases **86.5 kJ**. How many kilocalories of energy are released?

$$86.5 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 20.7 \text{ kcal}$$

Applications 3

04

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?

Applications 3

04

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?

$$\text{Unit X} = 0.1 \text{ cal}$$

$$1 \text{ cal} = 4.184 \text{ J}$$

$$X = (0.1 \text{ cal})(4.184 \text{ J/cal}) = 0.4184 \text{ J}$$

$$1 \text{ cal} = 0.001 \text{ Calorie}$$

$$X = (0.1 \text{ cal})(1 \text{ Cal}/1000 \text{ cal}) = 0.0001 \text{ Calorie}$$

Example 2

05

The temperature of a sample of iron with a mass of **10.0 g** changed from **50.4°C** to **25.0°C** with the release of **114 J**. What is the specific heat of iron?

Example 2

05

The temperature of a sample of iron with a mass of **10.0 g** changed from **50.4°C** to **25.0°C** with the release of **114 J**. What is the specific heat of iron?

Calculate ΔT .

$$\Delta T = 50.4^{\circ}\text{C} - 25.0^{\circ}\text{C} = 25.4^{\circ}\text{C}$$

Write the equation for calculating the quantity of heat.

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

State the equation for calculating heat.

$$\frac{\cancel{c} \times \cancel{m} \times \cancel{\Delta T}}{\cancel{m} \times \cancel{\Delta T}} = \frac{q}{m \times \Delta T}$$

Solve for c .

$$c = \frac{q}{m \times \Delta T}$$

$$c = \frac{114 \text{ J}}{(10.0 \text{ g})(25.4^{\circ}\text{C})}$$

Substitute $q = 114 \text{ J}$, $m = 10.0 \text{ g}$,
and $\Delta T = 25.4^{\circ}\text{C}$.

$$c = 0.449 \text{ J}/(\text{g}\cdot^{\circ}\text{C})$$

Multiply and divide numbers and units.

Example 2 – Page 09

1.1 Energy

Applications 4

o6

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](#).

Applications 4

o6

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](#).

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

$$q = 2.44 \text{ J/(g} \cdot ^\circ\text{C)} \times 34.4 \text{ g} \times 53.8^\circ\text{C} = 4.52 \times 10^3 \text{ J}$$

Applications 5

07

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](#).

Applications 5

07

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](#).

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

$$c = \frac{q}{m\Delta T} = \frac{(5696 \text{ J})}{(155 \text{ g})(40.0 - 25.0^\circ\text{C})} = 2.45 \text{ J}/(\text{g}\cdot^\circ\text{C})$$

The specific heat is very close to the value for ethanol.

Applications 6

o8

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? Refer to [Table 2](#).

Applications 6

o8

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? Refer to [Table 2](#).

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

$$\Delta T = \frac{q}{cm} = \frac{(276 \text{ J})}{(0.129 \text{ J/g}\cdot^{\circ}\text{C})(4.50 \text{ g})} = 475^{\circ}\text{C}$$

$$\Delta T = T_f - T_i = T_f - 25.0^{\circ}\text{C} = 475^{\circ}\text{C}$$

$$T_f = 5.00 \times 10^2^{\circ}\text{C}$$

Review 7

09

Chemical potential energy changes to heat in exothermic reactions and the heat is _____. In endothermic reactions, heat is _____ and changed to chemical potential energy.

- a) released – released
- b) released – absorbed
- c) absorbed – released
- d) absorbed – absorbed

Review 7

09

Chemical potential energy changes to heat in exothermic reactions and the heat is _____. In endothermic reactions, heat is _____ and changed to chemical potential energy.

- a) released – released
- b) released – absorbed**
- c) absorbed – released
- d) absorbed – absorbed

Review 8

10

Which of the following represent chemical potential energy is being changed to kinetic energy?

- a) two separated magnets
- b) an avalanche of snow
- c) books on library shelves
- d) a mountain stream
- e) a stock-car race
- f) separation of charge in a battery

Review 8

10

Which of the following represent chemical potential energy is being changed to kinetic energy?

- a) two separated magnets
- b) an avalanche of snow
- c) books on library shelves
- d) a mountain stream
- e) a stock-car race**
- f) separation of charge in a battery

Review 10

11

Calculate the amount of heat absorbed when **5.50 g** of aluminum is heated from **25.0 °C** to **95.0 °C**. The specific heat of aluminum is **0.897 J/(g-°C)**.

Review 10

11

Calculate the amount of heat absorbed when **5.50 g** of aluminum is heated from **25.0 °C** to **95.0 °C**. The specific heat of aluminum is **0.897 J/(g·°C)**.

$$q = cm\Delta T$$

$$q = (0.897 \text{ J/(g}\cdot\text{°C)})(5.50 \text{ g})(95.0\text{°C} - 25.0\text{°C})$$

$$q = 345 \text{ J}$$

Review 11

12

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. Refer to [Table 2](#).

Review 11

12

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. Refer to [Table 2](#).

The temperature change is inversely proportional to the specific heat:

aluminum, iron, silver, gold.

Mastering Concepts 53

13

_____ is a form of energy that flows from a warmer object to a cooler object.

- a) Heat
- b) Calorie
- c) Energy
- d) Joule

Mastering Concepts 53

13

_____ is a form of energy that flows from a warmer object to a cooler object.

- a) Heat
- b) Calorie
- c) Energy
- d) Joule

Mastering Concepts 53

14

_____ is a measure of the average kinetic energy of the particles in a sample of matter.

- a) Calorie
- b) Heat
- c) Temperature
- d) Joule

Mastering Concepts 53

14

_____ is a measure of the average kinetic energy of the particles in a sample of matter.

- a) Calorie
- b) Heat
- c) **Temperature**
- d) Joule

Mastering Concepts 54

15

How does the chemical potential energy of a system change during an endothermic reaction?

- a) increases
- b) decreases
- c) does not change
- d) depends on the system

Mastering Concepts 54

15

How does the chemical potential energy of a system change during an endothermic reaction?

- a) **increases**
- b) decreases
- c) does not change
- d) depends on the system

Mastering Concepts 55

16

During an avalanche, the _____ energy of snow at a higher altitude is converted to _____ energy as the snow cascades down a mountain.

- a) kinetic – potential
- b) chemical – kinetic
- c) potential – electrical
- d) potential – kinetic

Mastering Concepts 55

16

During an avalanche, the _____ energy of snow at a higher altitude is converted to _____ energy as the snow cascades down a mountain.

- a) kinetic – potential
- b) chemical – kinetic
- c) potential – electrical
- d) **potential – kinetic**

Mastering Concepts 56

17

How is the energy in gasoline converted and released when it burns in an automobile engine?

- a) Much is converted to work to move pistons in the engine; some is released as heat.
- b) Much is converted to heat to move pistons in the engine; some is released as work.
- c) Some is converted to heat to move pistons in the engine; much is released as work.
- d) Some is converted to work to move pistons in the engine; much is released as heat.

Mastering Concepts 56

17

How is the energy in gasoline converted and released when it burns in an automobile engine?

- a) Much is converted to work to move pistons in the engine; some is released as heat.
- b) Much is converted to heat to move pistons in the engine; some is released as work.
- c) Some is converted to heat to move pistons in the engine; much is released as work.
- d) **Some is converted to work to move pistons in the engine; much is released as heat.**

Mastering Concepts 57

18

How does the nutritional Calorie compare with the calorie?

- a) One nutritional Calorie equals 1 calorie.
- b) One nutritional Calorie equals 1000 calories.
- c) One nutritional Calorie equals 10 calories.
- d) One nutritional Calorie equals 0.001 calories.

Mastering Concepts 57

18

How does the nutritional Calorie compare with the calorie?

- a) One nutritional Calorie equals 1 calorie.
- b) One nutritional Calorie equals 1000 calories.**
- c) One nutritional Calorie equals 10 calories.
- d) One nutritional Calorie equals 0.001 calories.

Mastering Concepts 57

18

What is the relationship between the Calorie and a kilocalorie?

- a) One nutritional Calorie equals 10 kilocalorie.
- b) One nutritional Calorie equals 1000 kilocalorie.
- c) One nutritional Calorie equals 0.1 kilocalorie.
- d) One nutritional Calorie equals 1 kilocalorie.

Mastering Concepts 57

18

What is the relationship between the Calorie and a kilocalorie?

- a) One nutritional Calorie equals 10 kilocalorie.
- b) One nutritional Calorie equals 1000 kilocalorie.
- c) One nutritional Calorie equals 0.1 kilocalorie.
- d) One nutritional Calorie equals 1 kilocalorie.**

Mastering Concepts 58

19

What quantity has the units $\text{J}/(\text{g} \cdot ^\circ\text{C})$?

- a) Heat
- b) Specific heat
- c) Energy
- d) Temperature

Mastering Concepts 58

19

What quantity has the units $\text{J}/(\text{g} \cdot ^\circ\text{C})$?

- a) Heat
- b) Specific heat**
- c) Energy
- d) Temperature

Mastering Problems 62

20

A food item contains **124 nutritional Calories**. How many calories does the food item contain?

Mastering Problems 62

20

A food item contains **124 nutritional Calories**. How many calories does the food item contain?

$$124 \text{ Calories} \times \frac{1000 \text{ calories}}{1 \text{ Calorie}} = 124,000 \text{ calories}$$

Mastering Problems 63

21

How many joules are absorbed in a process that absorbs **0.5720 kcal**?

Mastering Problems 63

21

How many joules are absorbed in a process that absorbs **0.5720 kcal**?

$$0.5720 \text{ kcal} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} \times \frac{4.184 \text{ J}}{\text{cal}} = 2,393 \text{ J}$$

Mastering Problems 64

22

Ethanol is being used as an additive to gasoline. The combustion of **1 mol** of ethanol releases **1367 kJ** of energy. How many Calories are released?

Mastering Problems 64

22

Ethanol is being used as an additive to gasoline. The combustion of **1 mol** of ethanol releases **1367 kJ** of energy. How many Calories are released?

$$1367 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} \times \frac{1 \text{ Calorie}}{1000 \text{ cal}} \\ = 327 \text{ Calories}$$

Mastering Problems 65

23

To vaporize **2.00 g** of ammonia **656 calories** are required. How many kilojoules are required to vaporize the same mass of ammonia?

Mastering Problems 65

23

To vaporize **2.00 g** of ammonia **656 calories** are required. How many kilojoules are required to vaporize the same mass of ammonia?

$$656 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 2.74 \text{ kJ}$$

Mastering Problems 66

24

One mole of ethanol releases **326.7 Calories** of energy during combustion. How many kilojoules are released?

Mastering Problems 66

24

One mole of ethanol releases **326.7 Calories** of energy during combustion. How many kilojoules are released?

$$326.7 \text{ Cal} \times \frac{1000 \text{ cal}}{1 \text{ Cal}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 1367 \text{ kJ}$$

Mastering Problems 67

25

A **25.0-g** bolt made of an alloy absorbed **250 J** of heat as its temperature changed from **25.0 °C** to **78.0 °C**. What is the specific heat of the alloy?

Mastering Problems 67

25

A **25.0-g** bolt made of an alloy absorbed **250 J** of heat as its temperature changed from **25.0 °C** to **78.0 °C**. What is the specific heat of the alloy?

$$\Delta T = 78.0^{\circ}\text{C} - 25.0^{\circ}\text{C} = 53.0^{\circ}\text{C}$$

$$c = \frac{q}{m \times \Delta T} = \frac{250 \text{ J}}{25.0 \text{ g} \times 53.0^{\circ}\text{C}}$$

$$c = 0.189 \text{ J/g}\cdot^{\circ}\text{C}$$

Heat

Chapter 1 Lesson 2 – Revision Paper

Table 2

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	

Table 2 – Page 08

Example 3

1.2 Heat

Example 3

01

A piece of metal with a mass of **4.68 g** absorbs **256 J** of heat when its temperature increases by **182 °C**. What is the specific heat of the metal? Could the metal be one of the alkaline earth metals listed in [Table 2](#)?

Example 3

A piece of metal with a mass of **4.68 g** absorbs **256 J** of heat when its temperature increases by **182 °C**. What is the specific heat of the metal? Could the metal be one of the alkaline earth metals listed in [Table 2](#)?

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

State the equation for the quantity of heat, q .

$$c = \frac{q}{m \times \Delta T}$$

Solve for c .

$$c = \frac{256 \text{ J}}{(4.68 \text{ g})(182^\circ\text{C})} = 0.301 \text{ J}/(\text{g}\cdot^\circ\text{C})$$

Substitute $q = 256 \text{ J}$, $m = 4.68 \text{ g}$, and $\Delta T = 182^\circ\text{C}$.

Table 2 indicates that the metal could be strontium.

Applications 12

02

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?

Applications 12

02

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?

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

$$25.6 \text{ J} = c \times 90.0 \text{ g} \times 1.18^\circ\text{C}$$

$$c = 0.241 \text{ J/(g}\cdot^\circ\text{C)}$$

Applications 13

03

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?

Applications 13

03

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?

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

$$5650 \text{ J} = 4.184 \text{ J/(g} \cdot ^\circ\text{C)} \times m \times 26.6^\circ\text{C}$$

$$m = 50.8 \text{ g}$$

Applications 14

04

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

Applications 14

04

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

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

$$q = 0.803 \text{ J/(g}\cdot\text{°C)} \times 2.00 \times 10^3 \text{ g} \times 19.0\text{°C}$$

$$q = 30,500 \text{ J}$$

Applications 15

05

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

Applications 15

05

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

$$q = c \times m \times \Delta T = c \times m \times (T_f - T_i)$$

$$T_f = \frac{q}{cm} + T_i$$

$$T_f = \frac{-9750 \text{ J}}{(4.184 \text{ J/(g}\cdot\text{°C)})(335 \text{ g})} + 65.5^\circ\text{C}$$

$$T_f = 58.5^\circ\text{C}$$

Review 20

o6

The _____ is the particular part of the universe that contains the reaction or process that is being studied. The _____ are everything in the universe except the system. Thus the _____ is the system and its surroundings.

- a) system – universe – surroundings
- b) universe – surroundings – system
- c) universe – systems – surrounding
- d) system – surroundings – universe

Review 20

o6

The _____ is the particular part of the universe that contains the reaction or process that is being studied. The _____ are everything in the universe except the system. Thus the _____ is the system and its surroundings.

- a) system – universe – surroundings
- b) universe – surroundings – system
- c) universe – systems – surrounding
- d) **system – surroundings – universe**

Review 21

07

Calculate the specific heat of an unknown substance if a **2.50 g** sample releases **12.0 cal** as its temperature changes from **25.0 °C** to **20.0 °C**.

Review 21

07

Calculate the specific heat of an unknown substance if a **2.50 g** sample releases **12.0 cal** as its temperature changes from **25.0 °C** to **20.0 °C**.

$$q = cm\Delta T$$

$$c = \frac{q}{m\Delta T} = \frac{(-12 \text{ cal})(4.184 \text{ J/cal})}{(2.50 \text{ g})(-5.0^\circ\text{C})} = 4.02 \text{ J/(g}\cdot^\circ\text{C)}$$

Mastering Concepts 68

o8

The _____ is better insulated than a _____ , so that a minimal amount of heat is transferred into or out of the calorimeter.

- a) glass beaker – foam cup
- b) foam cup – glass beaker
- c) foam cup - bomb calorimeter
- d) glass beaker - bomb calorimeter

Mastering Concepts 68

o8

The _____ is better insulated than a _____ , so that a minimal amount of heat is transferred into or out of the calorimeter.

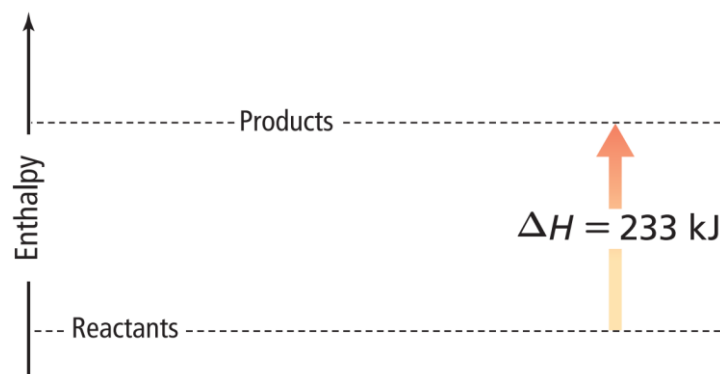
- a) glass beaker – foam cup
- b) foam cup – glass beaker**
- c) foam cup - bomb calorimeter
- d) glass beaker - bomb calorimeter

Mastering Concepts 69

09

Is the reaction shown in the figure shown endothermic or exothermic?

- a) Endothermic, because the enthalpy is positive
- b) Exothermic, because the enthalpy is positive
- c) Endothermic, because the enthalpy is negative
- d) Exothermic, because the enthalpy is negative



Mastering Concepts 69 – Page 40

1.2 Heat

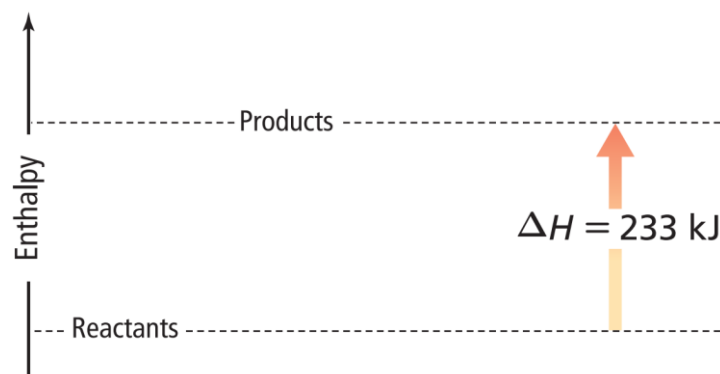


Mastering Concepts 69

09

Is the reaction shown in the figure shown endothermic or exothermic?

- a) Endothermic, because the enthalpy is positive
- b) Exothermic, because the enthalpy is positive
- c) Endothermic, because the enthalpy is negative
- d) Exothermic, because the enthalpy is negative



Mastering Concepts 69 – Page 40

1.2 Heat



Mastering Concepts 70

10

During this exothermic heat-pack reaction, heat produced by the reaction flows from the heat pack _____ to your cold hands _____.

- a) system – surroundings
- b) universe – system
- c) system – part of the surroundings
- d) surroundings – system

Mastering Concepts 70

10

During this exothermic heat-pack reaction, heat produced by the reaction flows from the heat pack _____ to your cold hands _____.

- a) system – surroundings
- b) universe – system
- c) **system – part of the surroundings**
- d) surroundings – system

Mastering Concepts 71

11

Under what condition is the heat (q) evolved or absorbed in a chemical reaction equal to a change in enthalpy (ΔH)?

- a) at constant pressure
- b) at constant temperature
- c) at constant energy
- d) at variable pressure

Mastering Concepts 71

11

Under what condition is the heat (q) evolved or absorbed in a chemical reaction equal to a change in enthalpy (ΔH)?

- a) at constant pressure
- b) at constant temperature
- c) at constant energy
- d) at variable pressure

Mastering Concepts 72

12

The enthalpy change for a reaction, ΔH , is **negative**. What does this indicate about the chemical potential energy of the system before and after the reaction?

- a) After > Before
- b) Before > After
- c) Before < After
- d) Before = After

Mastering Concepts 72

12

The enthalpy change for a reaction, ΔH , is **negative**. What does this indicate about the chemical potential energy of the system before and after the reaction?

- a) After > Before
- b) Before > After**
- c) Before < After
- d) Before = After

Mastering Concepts 73

13

What is the sign of ΔH for an exothermic reaction? An endothermic reaction?

- a) Exothermic (+) – Endothermic (-)
- b) Exothermic (+) – Endothermic (+)
- c) Exothermic (-) – Endothermic (-)
- d) Exothermic (-) – Endothermic (+)

Mastering Concepts 73

13

What is the sign of ΔH for an exothermic reaction? An endothermic reaction?

- a) Exothermic (+) – Endothermic (-)
- b) Exothermic (+) – Endothermic (+)
- c) Exothermic (-) – Endothermic (-)
- d) **Exothermic (-) – Endothermic (+)**

Mastering Problems 74

14

How many joules of heat are lost by **3580 kg** of granite as it cools from **41.2°C** to **-12.9°C**? The specific heat of granite is **0.803 J/(g·°C)**.

Mastering Problems 74

14

How many joules of heat are lost by **3580 kg** of granite as it cools from **41.2°C** to **-12.9°C**? The specific heat of granite is **0.803 J/(g·°C)**.

$$\Delta T = 41.2^{\circ}\text{C} - (-12.9^{\circ}\text{C}) = 54.1^{\circ}\text{C}$$

$$q_{\text{granite}} = [0.803 \text{ J/(g}\cdot^{\circ}\text{C)}](3.58 \times 10^6 \text{ g})(54.1^{\circ}\text{C})$$

$$q_{\text{granite}} = 1.56 \times 10^8 \text{ J}$$

Mastering Problems 75

15

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**.

Mastering Problems 75

15

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**.

Change the dimensions of the pool's water from meters to centimeters.

$$20.0 \text{ m} = 2.00 \times 10^3 \text{ cm}; 12.5 \text{ m} = 1.25 \times 10^3 \text{ cm}; \\ 3.75 \text{ m} = 3.75 \times 10^2 \text{ cm}$$

$$\text{volume of water} = (2.00 \times 10^3 \text{ cm})(1.25 \times 10^3 \text{ cm}) \\ (3.75 \times 10^2 \text{ cm}) = 9.38 \times 10^8 \text{ cm}^3 \\ = 9.38 \times 10^8 \text{ mL}$$

$$\text{mass of water} = (9.38 \times 10^8 \text{ mL})(1.000 \text{ g/mL}) \\ = 9.38 \times 10^8 \text{ g}$$

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

$$\Delta T = (29.0^\circ\text{C} - 18.4^\circ\text{C}) = 10.6^\circ\text{C}$$

$$q = [4.184 \text{ J/(g}\cdot^\circ\text{C)}](9.38 \times 10^8 \text{ g})(10.6^\circ\text{C}) \\ = 4.16 \times 10^{10} \text{ J}$$

Mastering Problems 76

16

How much heat is absorbed by a **44.7 g** piece of lead when its temperature increases by **65.4°C**? The specific heat of lead is **0.129 J/(g·°C)**.

Mastering Problems 76

16

How much heat is absorbed by a **44.7 g** piece of lead when its temperature increases by **65.4°C**? The specific heat of lead is **0.129 J/(g·°C)**.

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

$$q = 0.129 \text{ J/(g}\cdot\text{°C)} \times 44.7 \text{ g} \times 65.4\text{°C} = 377 \text{ J}$$

Mastering Problems 77

17

When **10.2 g** of canola oil at **25.0 °C** is placed in a wok, **3.34 kJ** of heat is required to heat it to a temperature of **196.4°C**. What is the specific heat of the canola oil?

Mastering Problems 77

17

When **10.2 g** of canola oil at **25.0 °C** is placed in a wok, **3.34 kJ** of heat is required to heat it to a temperature of **196.4°C**. What is the specific heat of the canola oil?

$$3.34 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 3340 \text{ J}$$

$$\Delta T = T_f - T_i = 196.4^\circ\text{C} - 25.0^\circ\text{C} = 171.4^\circ\text{C}$$

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

$$c = \frac{q}{m \times \Delta T} = \frac{3340 \text{ J}}{10.2 \text{ g} \times 171.4^\circ\text{C}} = 1.91 \text{ J}/(\text{g}\cdot^\circ\text{C})$$

Mastering Problems 78

18

When a **58.8 g** piece of hot alloy is placed in **125 g** of cold water in a calorimeter, the temperature of the alloy decreases by **106.1 °C**, while the temperature of the water increases by **10.5 °C**. What is the specific heat of the alloy?

Mastering Problems 78

18

When a **58.8 g** piece of hot alloy is placed in **125 g** of cold water in a calorimeter, the temperature of the alloy decreases by **106.1 °C**, while the temperature of the water increases by **10.5 °C**. What is the specific heat of the alloy?

$$q = c \times m \times \Delta T; q_{\text{water}} = q_{\text{alloy}}$$

$$4.184 \text{ J/(g}\cdot\text{°C)} \times 125 \text{ g} \times 10.5\text{°C} \\ = c_{\text{alloy}} \times 58.8 \text{ g} \times 106.1\text{°C}$$

$$c_{\text{alloy}} = \frac{(4.184 \text{ J/g}\cdot\text{°C})(125 \text{ g})(10.5\text{°C})}{(58.8 \text{ g})(106.1\text{°C})}$$

$$c_{\text{alloy}} = 0.880 \text{ J/(g}\cdot\text{°C)}$$

THINK CRITICALLY 110

19

Suppose that two pieces of iron, one with a mass exactly twice the mass of the other, are placed in an insulated calorimeter. If the original temperatures of the larger piece and the smaller piece are **90.0 °C** and **50.0 °C**, respectively, what is the temperature of the two pieces when thermal equilibrium has been established? The specific heat of iron is **0.449 J/(g·°C)**.

THINK CRITICALLY 110

19

Suppose that two pieces of iron, one with a mass exactly twice the mass of the other, are placed in an insulated calorimeter. If the original temperatures of the larger piece and the smaller piece are **90.0 °C** and **50.0 °C**, respectively, what is the temperature of the two pieces when thermal equilibrium has been established? The specific heat of iron is **0.449 J/(g·°C)**.

Let subscript 1 refer to the smaller, cooler piece.

Let subscript 2 refer to the larger, hotter piece.

Heat lost by the hotter piece = heat gained by cooler piece:

$$-q_1 = q_2$$

$$-cm_1(T_1 - T_f) = cm_2(T_2 - T_f); T_f = \text{final temperature}$$

Eliminate the specific heat c from this equation:

From the problem statement: $m_2 = 2m_1$:

$$-m_1(T_1 - T_f) = 2m_1(T_2 - T_f)$$

Eliminate mass m_1 from this equation:

$$-(T_1 - T_f) = 2(T_2 - T_f)$$

Solve for the unknown T_f :

$$T_f = \frac{1}{3}(T_1 + 2T_2) = \frac{1}{3}(50^\circ\text{C} + 2(90^\circ\text{C})) = 76.7^\circ\text{C}$$

The result is a mass-weighted average of the two temperatures.

THINK CRITICALLY 110 – Page 42

1.2 Heat

SAT SUBJECT TEST 14

20

The specific heat of ethanol is $2.44 \text{ J/g}\cdot^{\circ}\text{C}$. How many kilojoules of energy are required to heat 50.0 g of ethanol from -20.0°C to 68.0°C ?

- a) 10.7 kJ
- b) 1.22 kJ
- c) 8.30 kJ
- d) 5.86 kJ
- e) 2.44 kJ

SAT SUBJECT TEST 14

20

The specific heat of ethanol is **2.44 J/g·°C**. How many kilojoules of energy are required to heat **50.0 g** of ethanol from **-20.0 °C** to **68.0 °C**?

a) 10.7 kJ

b) 1.22 kJ

c) 8.30 kJ

d) 5.86 kJ

e) 2.44 kJ

$$q = cm\Delta T = (2.44 \text{ J/(g}\cdot\text{°C)}) \times (50.0 \text{ g}) \times (88.0\text{°C}) \times \frac{1\text{kJ}}{1000 \text{ J}} = 10.7 \text{ kJ}$$

SAT SUBJECT TEST 15

21

If **3.00 g** of aluminum foil, placed in an oven and heated from **20.0 °C** to **662.0 °C**, absorbs **1728 J** of heat, what is the specific heat of aluminum?

- a) $0.131 \text{ J/g}\cdot^{\circ}\text{C}$
- b) $2.61 \text{ J/g}\cdot^{\circ}\text{C}$
- c) $0.870 \text{ J/g}\cdot^{\circ}\text{C}$
- d) $0.261 \text{ J/g}\cdot^{\circ}\text{C}$
- e) $0.897 \text{ J/g}\cdot^{\circ}\text{C}$

SAT SUBJECT TEST 15

21

If **3.00 g** of aluminum foil, placed in an oven and heated from **20.0 °C** to **662.0 °C**, absorbs **1728 J** of heat, what is the specific heat of aluminum?

a) $0.131 \text{ J/g}\cdot^\circ\text{C}$

$$q = cm\Delta T$$

b) $2.61 \text{ J/g}\cdot^\circ\text{C}$

$$c = \frac{q}{m\Delta T} = \frac{(1728 \text{ J})}{(3.00 \text{ g})(642.0^\circ\text{C})} = 0.897 \text{ J/(g}\cdot^\circ\text{C)}$$

c) $0.870 \text{ J/g}\cdot^\circ\text{C}$

d) $0.261 \text{ J/g}\cdot^\circ\text{C}$

e) $0.897 \text{ J/g}\cdot^\circ\text{C}$

Thermochemical Equations

Chapter 1 Lesson 3 – Revision Paper



مكتبة الفكر

Example 4

01

How much heat is evolved when **54.0 g** glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) is burned according to this equation?



Example 4 – Page 20

Example 4

01

How much heat is evolved when **54.0 g** glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) is burned according to this equation?



Convert grams of $\text{C}_6\text{H}_{12}\text{O}_6$ to moles of $\text{C}_6\text{H}_{12}\text{O}_6$.

$$54.0 \text{ g } \cancel{\text{C}_6\text{H}_{12}\text{O}_6} \times \frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.18 \text{ g } \cancel{\text{C}_6\text{H}_{12}\text{O}_6}} = 0.300 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6$$

Multiply by the inverse of molar mass, $\frac{1 \text{ mol}}{180.18 \text{ g}}$.

Multiply moles of $\text{C}_6\text{H}_{12}\text{O}_6$ by the enthalpy of combustion, ΔH_{comb} .

$$0.300 \text{ mol } \cancel{\text{C}_6\text{H}_{12}\text{O}_6} \times \frac{2808 \text{ kJ}}{1 \text{ mol } \cancel{\text{C}_6\text{H}_{12}\text{O}_6}} = 842 \text{ kJ}$$

Multiply moles of glucose by $\frac{2808 \text{ kJ}}{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}$.

Example 4 – Page 20



APPLICATIONS 23

02

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

Table 4 Standard Enthalpies of Vaporization and Fusion			
Substance	Formula	$\Delta H_{\text{vap}}^{\circ}$ (kJ/mol)	$\Delta H_{\text{fus}}^{\circ}$ (kJ/mol)
Water	H ₂ O	40.7	6.01
Ethanol	C ₂ H ₅ OH	38.6	4.94
Methanol	CH ₃ OH	35.2	3.22
Acetic acid	CH ₃ COOH	23.4	11.7
Ammonia	NH ₃	23.3	5.66

APPLICATIONS 23

02

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

$$25.7 \text{ g } \cancel{\text{CH}_3\text{OH}} \times \frac{1 \text{ mol } \cancel{\text{CH}_3\text{OH}}}{32.04 \text{ g } \cancel{\text{CH}_3\text{OH}}} \times \frac{3.22 \text{ kJ}}{1 \text{ mol } \cancel{\text{CH}_3\text{OH}}} = 2.58 \text{ kJ}$$

Table 4 Standard Enthalpies of Vaporization and Fusion			
Substance	Formula	$\Delta H_{\text{vap}}^\circ$ (kJ/mol)	$\Delta H_{\text{fus}}^\circ$ (kJ/mol)
Water	H ₂ O	40.7	6.01
Ethanol	C ₂ H ₅ OH	38.6	4.94
Methanol	CH ₃ OH	35.2	3.22
Acetic acid	CH ₃ COOH	23.4	11.7
Ammonia	NH ₃	23.3	5.66

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} .

Table 4 Standard Enthalpies of Vaporization and Fusion			
Substance	Formula	$\Delta H_{\text{vap}}^{\circ}$ (kJ/mol)	$\Delta H_{\text{fus}}^{\circ}$ (kJ/mol)
Water	H ₂ O	40.7	6.01
Ethanol	C ₂ H ₅ OH	38.6	4.94
Methanol	CH ₃ OH	35.2	3.22
Acetic acid	CH ₃ COOH	23.4	11.7
Ammonia	NH ₃	23.3	5.66

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} .

$$275 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \times \frac{23.3 \text{ kJ}}{1 \text{ mol NH}_3} = 376 \text{ kJ}$$

Table 4 Standard Enthalpies of Vaporization and Fusion			
Substance	Formula	$\Delta H_{\text{vap}}^{\circ}$ (kJ/mol)	$\Delta H_{\text{fus}}^{\circ}$ (kJ/mol)
Water	H ₂ O	40.7	6.01
Ethanol	C ₂ H ₅ OH	38.6	4.94
Methanol	CH ₃ OH	35.2	3.22
Acetic acid	CH ₃ COOH	23.4	11.7
Ammonia	NH ₃	23.3	5.66

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

Table 3 Standard Enthalpies of Combustion		
Substance	Formula	$\Delta H^\circ_{\text{comb}}(\text{kJ/mol})$
Sucrose (table sugar)	$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$	−5644
Octane (a component of gasoline)	$\text{C}_8\text{H}_{18}(\text{l})$	−5471
Glucose (a simple sugar found in fruit)	$\text{C}_6\text{H}_{12}\text{O}_6(\text{s})$	−2808
Propane (a gaseous fuel)	$\text{C}_3\text{H}_8(\text{g})$	−2219
Methane (a gaseous fuel)	$\text{CH}_4(\text{g})$	−891

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

$$12,880 \text{ kJ} = m \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{891 \text{ kJ}}{1 \text{ mol CH}_4}$$

$$m = 12,880 \text{ kJ} \times \frac{16.04 \text{ g CH}_4}{1 \text{ mol CH}_4} \times \frac{1 \text{ mol CH}_4}{891 \text{ kJ}}$$

$$m = 232 \text{ g CH}_4$$

Table 3 Standard Enthalpies of Combustion

Substance	Formula	$\Delta H^\circ_{\text{comb}}(\text{kJ/mol})$
Sucrose (table sugar)	$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$	−5644
Octane (a component of gasoline)	$\text{C}_8\text{H}_{18}(\text{l})$	−5471
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Review 26

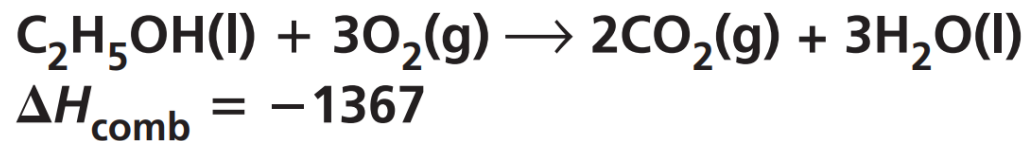
05

Write a complete thermochemical equation for the combustion of ethanol ($\text{C}_2\text{H}_5\text{OH}$). $\Delta H_{\text{comb}} = -1367 \text{ kJ/mol}$

Review 26

05

Write a complete thermochemical equation for the combustion of ethanol ($\text{C}_2\text{H}_5\text{OH}$). $\Delta H_{\text{comb}} = -1367 \text{ kJ/mol}$



Review 27

o6

Which of the following processes are **exothermic**? **Endothermic**?

- a. $\text{C}_2\text{H}_5\text{OH}(\text{l}) \rightarrow \text{C}_2\text{H}_5\text{OH}(\text{g})$
- b. $\text{Br}_2(\text{l}) \rightarrow \text{Br}_2(\text{s})$
- c. $\text{C}_5\text{H}_{12}(\text{g}) + 8\text{O}_2(\text{g}) \rightarrow 5\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l})$
- d. $\text{NH}_3(\text{g}) \rightarrow \text{NH}_3(\text{l})$
- e. $\text{NaCl}(\text{s}) \rightarrow \text{NaCl}(\text{l})$

Review 27

o6

Which of the following processes are **exothermic**? **Endothermic**?

- a.** $\text{C}_2\text{H}_5\text{OH}(\text{l}) \rightarrow \text{C}_2\text{H}_5\text{OH}(\text{g})$
- b.** $\text{Br}_2(\text{l}) \rightarrow \text{Br}_2(\text{s})$
- c.** $\text{C}_5\text{H}_{12}(\text{g}) + 8\text{O}_2(\text{g}) \rightarrow 5\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l})$
- d.** $\text{NH}_3(\text{g}) \rightarrow \text{NH}_3(\text{l})$
- e.** $\text{NaCl}(\text{s}) \rightarrow \text{NaCl}(\text{l})$

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

Review 28

07

Calculate the heat released in freezing 0.250 mol water.

Review 28

07

Calculate the heat released in freezing 0.250 mol water.

$$q = n(-\Delta H^\circ_{\text{fus}})$$

$$q = 0.250 \text{ mol} \times -6.01 \text{ kJ/mol}$$

$$q = -1.50 \text{ kJ}$$

Review 29

o8

How much heat is released by the combustion of **206 g** of hydrogen gas?

$$\Delta H_{\text{comb}} = \mathbf{-286 \text{ kJ/mol}}$$

Review 29

o8

How much heat is released by the combustion of **206 g** of hydrogen gas?

$$\Delta H_{\text{comb}} = -286 \text{ kJ/mol}$$

The molar mass of hydrogen is 2.01 g/mol.

$$206 \text{ g} \times \frac{1 \text{ mol}}{2.01 \text{ g}} \times \frac{286 \text{ kJ}}{1 \text{ mol}} = 29,300 \text{ kJ}$$

Review 30

09

The molar heat of vaporization of ammonia is **23.3 kJ/mol**. What is the molar heat of condensation of ammonia?

Review 30

09

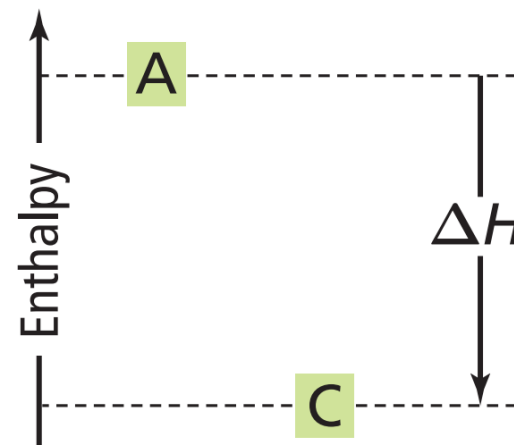
The molar heat of vaporization of ammonia is **23.3 kJ/mol**. What is the molar heat of condensation of ammonia?

-23.3 kJ/mol

Review 31

10

The reaction $A \rightarrow C$ is shown in the enthalpy diagram at right. Is the reaction **exothermic** or **endothermic**? Explain your answer.

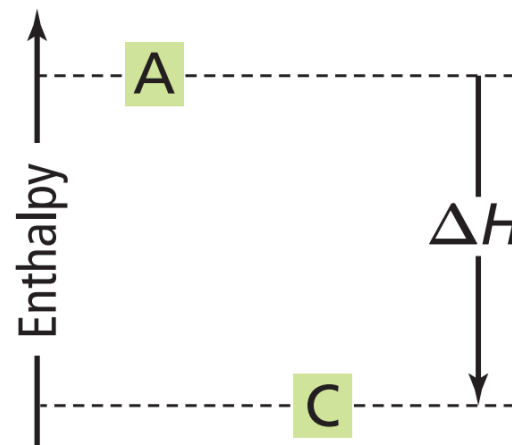


Review 31

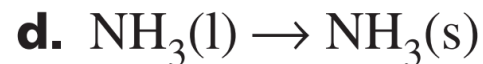
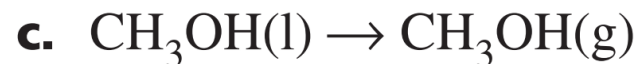
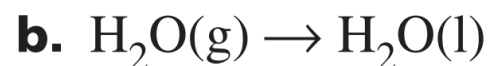
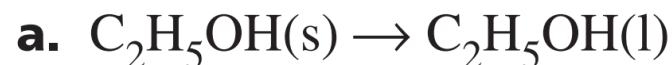
10

The reaction $A \rightarrow C$ is shown in the enthalpy diagram at right. Is the reaction **exothermic** or **endothermic**? Explain your answer.

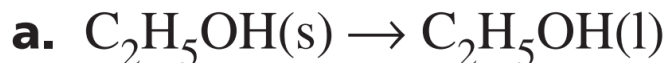
The reaction is **exothermic** because the product (C) has a lower enthalpy than the reactant (A).



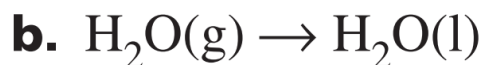
Write the sign of ΔH_{system} for each of the following changes in physical state.



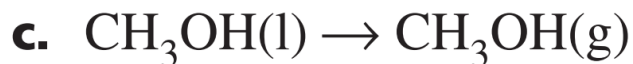
Write the sign of ΔH_{system} for each of the following changes in physical state.



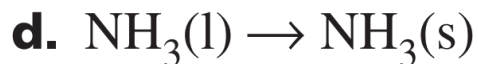
ΔH_{system} is positive.



ΔH_{system} is negative.



ΔH_{system} is positive.



ΔH_{system} is negative.

Write the thermochemical equation for the combustion of methane. Refer to **Table 3**.

Table 3 Standard Enthalpies of Combustion		
Substance	Formula	$\Delta H_{\text{comb}}^{\circ}(\text{kJ/mol})$
Sucrose (table sugar)	$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$	−5644
Octane (a component of gasoline)	$\text{C}_8\text{H}_{18}(\text{l})$	−5471
Glucose (a simple sugar found in fruit)	$\text{C}_6\text{H}_{12}\text{O}_6(\text{s})$	−2808
Propane (a gaseous fuel)	$\text{C}_3\text{H}_8(\text{g})$	−2219
Methane (a gaseous fuel)	$\text{CH}_4(\text{g})$	−891

Write the thermochemical equation for the combustion of methane. Refer to **Table 3**.

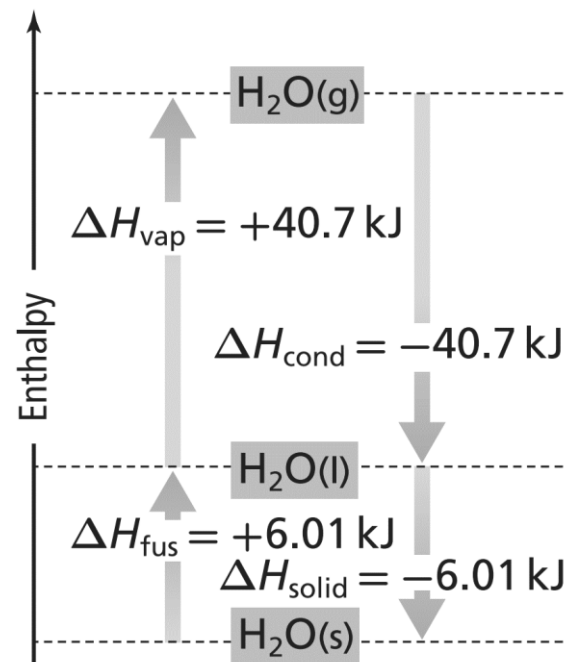


Table 3 Standard Enthalpies of Combustion		
Substance	Formula	$\Delta H^\circ_{\text{comb}}(\text{kJ/mol})$
Sucrose (table sugar)	$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$	-5644
Octane (a component of gasoline)	$\text{C}_8\text{H}_{18}(\text{l})$	-5471
Glucose (a simple sugar found in fruit)	$\text{C}_6\text{H}_{12}\text{O}_6(\text{s})$	-2808
Propane (a gaseous fuel)	$\text{C}_3\text{H}_8(\text{g})$	-2219
Methane (a gaseous fuel)	$\text{CH}_4(\text{g})$	-891

Mastering Problems 83

13

Use information from **Figure 24** to calculate how much heat is required to vaporize **4.33 mol** of water at **100°C**.



■ Figure 24



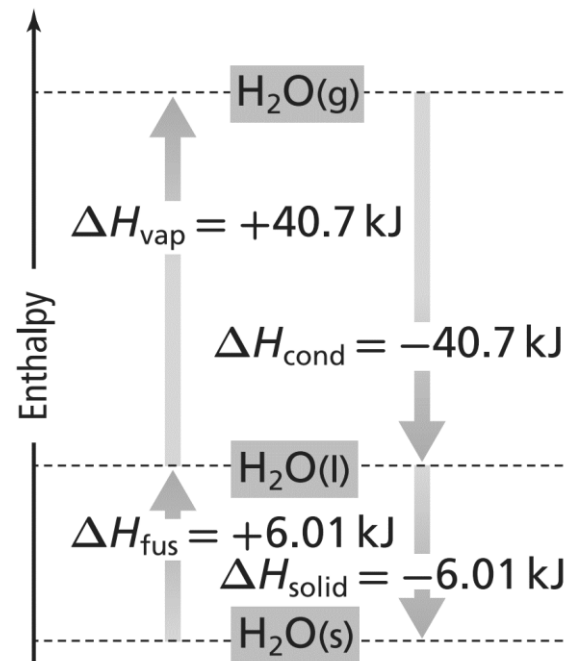
Mastering Problems 83

13

Use information from **Figure 24** to calculate how much heat is required to vaporize **4.33 mol** of water at **100°C**.

$$q = \text{mol} \times \Delta H_{\text{vap}}$$

$$q = 4.33 \text{ mol} \times 40.7 \text{ kJ/mol} = 176 \text{ kJ}$$



■ Figure 24

Mastering Problems 84

14

Water is sprayed on oranges during a frosty night. If an average of **11.8 g** of water freezes on each orange, how much heat is released?

Mastering Problems 84

14

Water is sprayed on oranges during a frosty night. If an average of **11.8 g** of water freezes on each orange, how much heat is released?

$$11.8 \text{ g H}_2\text{O} \times \frac{1 \text{ mole H}_2\text{O}}{18.0 \text{ g}} = 0.656 \text{ mol H}_2\text{O}$$

$$q = \text{mol} \times \Delta H_{\text{solid}}$$

$$q = 0.656 \text{ mol} \times (-6.01 \text{ kJ/mol}) = -3.94 \text{ kJ}$$



Mastering Problems 85

15

What mass of propane (C_3H_8) must be burned in a barbecue grill to release **4560 kJ** of heat? The ΔH_{comb} of propane is **-2219 kJ/mol**.

Mastering Problems 85

15

What mass of propane (C_3H_8) must be burned in a barbecue grill to release **4560 kJ** of heat? The ΔH_{comb} of propane is **-2219 kJ/mol**.

$$q = \text{mol} \times \Delta H_{\text{comb}}$$

$$\text{moles of propane} = \frac{4560 \text{ kJ}}{2219 \text{ kJ/mol}} = 2.055 \text{ mol}$$

$$2.055 \text{ mol C}_3\text{H}_8 \times 44.09 \text{ g C}_3\text{H}_8/\text{mol C}_3\text{H}_8 = 90.60 \text{ g}$$

Mastering Problems 86

16

How much heat is liberated when **5.00 kg** of coal is burned if the coal is **96.2%** carbon by mass and the other materials in the coal do not react? ΔH_{comb} of carbon is **-394 kJ/mol**.

Mastering Problems 86

16

How much heat is liberated when **5.00 kg** of coal is burned if the coal is **96.2%** carbon by mass and the other materials in the coal do not react? ΔH_{comb} of carbon is **-394 kJ/mol**.

$$m_{\text{carbon}} = m_{\text{coal}} \times 0.962 \times \frac{1000 \text{ g}}{1 \text{ kg}}$$
$$= (5.00 \text{ kg})(0.962)\left(\frac{1000 \text{ g}}{1 \text{ kg}}\right) = 4810 \text{ g}$$

$$\text{mol C} = 4810 \text{ g C} \times \frac{1 \text{ mol}}{12.0 \text{ g C}} = 401 \text{ mol C}$$

$$q = \text{mol} \times \Delta H_{\text{comb}}$$

$$q = 401 \text{ mol C} \times (-394 \text{ kJ/mol C}) = -158,000 \text{ kJ}$$

Mastering Problems 86 – Page 41

1.3 Thermochemical Equations

Mastering Problems 87

17

How much heat is evolved when **1255 g** of water condenses to a liquid at **100°C**?

Mastering Problems 87

17

How much heat is evolved when **1255 g** of water condenses to a liquid at **100°C**?

$$1255 \text{ g} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{40.7 \text{ kJ}}{1 \text{ mol}} = 2830 \text{ kJ}$$

Mastering Problems 88

18

A sample of ammonia ($\Delta H_{\text{solid}} = -5.66 \text{ kJ/mol}$) liberates **5.66 kJ** of heat as it solidifies at its melting point. What is the mass of the sample?

Mastering Problems 88

18

A sample of ammonia ($\Delta H_{\text{solid}} = -5.66 \text{ kJ/mol}$) liberates **5.66 kJ** of heat as it solidifies at its melting point. What is the mass of the sample?

Mass = mass of 1 mol ammonia = 17.03 g

Calculating Enthalpy Change

Chapter 1
Lesson 4 – Revision Paper



مكتبة الفكر

Table R-11 – Heat of Formation Values

Table R-11 Heat of Formation Values							
ΔH_f° (kJ/mol) (concentration of aqueous solutions is 1M)							
Substance	ΔH_f°	Substance	ΔH_f°	Substance	ΔH_f°	Substance	ΔH_f°
Ag(s)	0	CsCl(s)	-443.0	H ₃ PO ₄ (aq)	-1271.7	NaBr(s)	-361.1
AgCl(s)	-127.0	Cs ₂ SO ₄ (s)	-1443.0	H ₂ S(g)	-20.6	NaCl(s)	-411.2
AgCN(s)	146.0	CuI(s)	-67.8	H ₂ SO ₃ (aq)	-608.8	NaHCO ₃ (s)	-950.8
Al ₂ O ₃	-1675.7	CuS(s)	-53.1	H ₂ SO ₄ (aq)	-814.0	NaNO ₃ (s)	-467.9
BaCl ₂ (aq)	-855.0	Cu ₂ S(s)	-79.5	HgCl ₂ (s)	-224.3	NaOH(s)	-425.8
BaSO ₄	-1473.2	CuSO ₄ (s)	-771.4	Hg ₂ Cl ₂ (s)	-265.4	Na ₂ CO ₃ (s)	-1130.7
BeO(s)	-609.4	F ₂ (g)	0	Hg ₂ SO ₄ (s)	-743.1	Na ₂ S(s)	-364.8
BiCl ₃ (s)	-379.1	FeCl ₃ (s)	-399.49	I ₂ (s)	0	Na ₂ SO ₄ (s)	-1387.1
Bi ₂ S ₃ (s)	-143.1	FeO(s)	-272.0	K(s)	0	NH ₄ Cl(s)	-314.4
Br ₂	0	FeS(s)	-100.0	KBr(s)	-393.8	O ₂ (g)	0
CCl ₄ (l)	-128.2	Fe ₂ O ₃ (s)	-824.2	KMnO ₄ (s)	-837.2	P ₄ O ₆ (s)	-1640.1
CH ₄ (g)	-74.6	Fe ₃ O ₄ (s)	-1118.4	KOH	-424.6	P ₄ O ₁₀ (s)	-2984.0
C ₂ H ₂ (g)	227.4	H(g)	218.0	LiBr(s)	-351.2	PbBr ₂ (s)	-278.7
C ₂ H ₄ (g)	52.4	H ₂ (g)	0	LiOH(s)	-487.5	PbCl ₂ (s)	-359.4
C ₂ H 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)	-395.7
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 ₄ Br(s)	-270.8	UCl ₆ (s)	-1092.0
Cl ₂ (g)	0	H ₂ O(l)	-285.8	NO(g)	91.3	Zn(s)	0
Co ₃ O ₄ (s)	-891.0	H ₂ O(g)	-241.8	NO ₂ (g)	33.2	ZnCl ₂ (aq)	-415.1
CoO(s)	-237.9	H ₂ O ₂ (l)	-187.8	N ₂ O(g)	81.6	ZnO(s)	-350.5
Cr ₂ O ₃ (s)	-1139.7	H ₃ PO ₂ (l)	-595.4	Na(s)	0	ZnSO ₄ (s)	-982.8

Table R-11 – Heat of Formation Values – RT-8

[Example 6](#) - [Q.35](#) - [Q.36](#) - [Q.92](#)

1.4 Calculating Enthalpy Change

Figure 13

01

Determine the enthalpy change for the decomposition of SO_3 to S and O_2 .

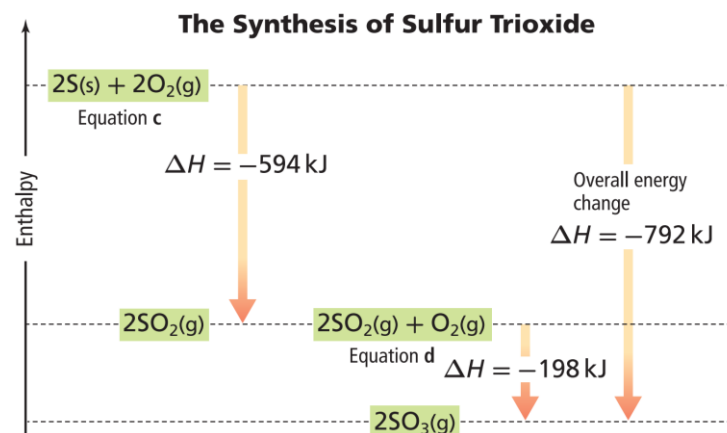


Figure 13 – Page 23

1.4 Calculating Enthalpy Change



Figure 13

01

Determine the enthalpy change for the decomposition of SO_3 to S and O_2 .

$+792 \text{ kJ}$

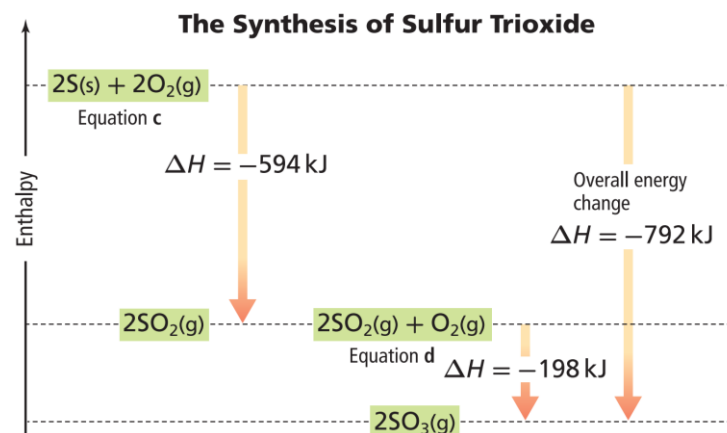


Figure 13 – Page 23

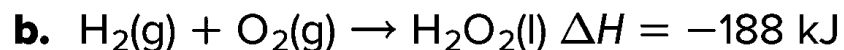
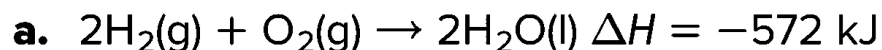
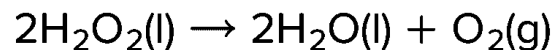
1.4 Calculating Enthalpy Change



EXAMPLE 5

02

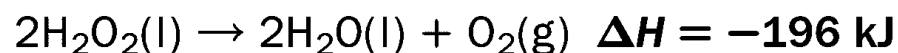
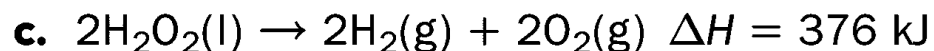
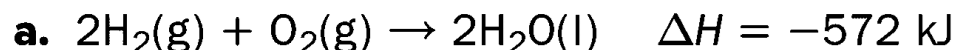
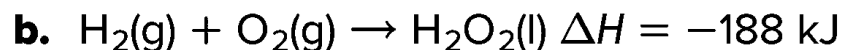
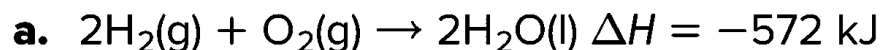
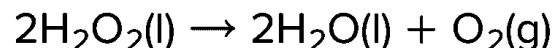
Use thermochemical Equations **a** and **b** below to determine ΔH for the decomposition of hydrogen peroxide (**H₂O₂**).



EXAMPLE 5

02

Use thermochemical Equations **a** and **b** below to determine ΔH for the decomposition of hydrogen peroxide (**H₂O₂**).



EXAMPLE 5 – Page 24

1.4 Calculating Enthalpy Change

APPLICATIONS 32

03

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



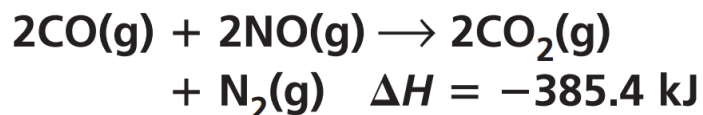
APPLICATIONS 32

03

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



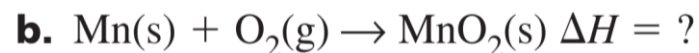
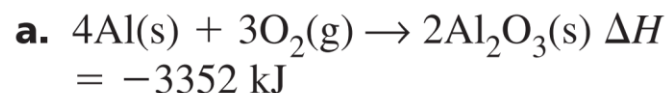
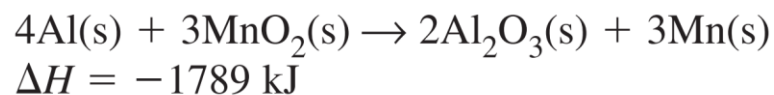
Add the first equation to the second equation reversed.



APPLICATIONS 33

04

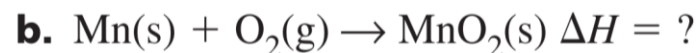
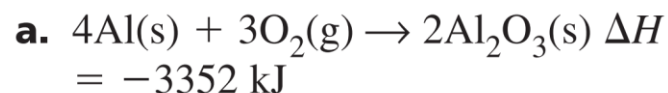
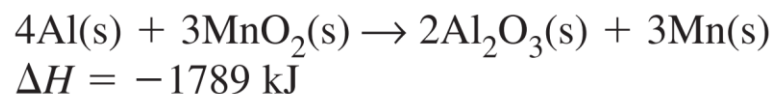
ΔH for the following reaction is **-1789 kJ**. Use this and Equation **a** to determine ΔH for Equation **b**.



APPLICATIONS 33

04

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



Let $x = \Delta H$ for equation **b**.

Add the Equation **a** to Equation **b** reversed and tripled.

$$\Delta H = -3352 - 3x \text{ kJ}$$

$$-3352 - 3x \text{ kJ} = -1789 \text{ kJ}$$

Because the direction of Equation **b** was changed, ΔH for equation **b** = x =

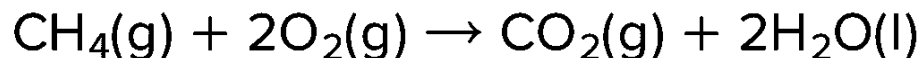
$$\frac{-3352 + 1789}{3} = -521 \text{ kJ}$$



EXAMPLE 6

05

Use standard enthalpies of formation to calculate $\Delta H_{\text{rxn}}^\circ$ for the combustion of **methane**. ([Table R-11](#))



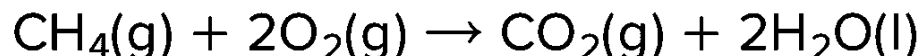
EXAMPLE 6 – Page 28

1.4 Calculating Enthalpy Change

EXAMPLE 6

05

Use standard enthalpies of formation to calculate $\Delta H_{\text{rxn}}^{\circ}$ for the combustion of **methane**. ([Table R-11](#))



Use the formula $\Delta H_{\text{rxn}}^{\circ} = \sum \Delta H_{\text{f}}^{\circ}(\text{products}) - \sum \Delta H_{\text{f}}^{\circ}(\text{reactants})$.

$$\Delta H_{\text{rxn}}^{\circ} = [\Delta H_{\text{f}}^{\circ}(\text{CO}_2) + (2)\Delta H_{\text{f}}^{\circ}(\text{H}_2\text{O})] - [\Delta H_{\text{f}}^{\circ}(\text{CH}_4) + (2)\Delta H_{\text{f}}^{\circ}(\text{O}_2)]$$

$$\Delta H_{\text{rxn}}^{\circ} = [(-394 \text{ kJ}) + (2)(-286 \text{ kJ})] - [(-75 \text{ kJ}) + (2)(0.0 \text{ kJ})]$$

$$\Delta H_{\text{rxn}}^{\circ} = [-966 \text{ kJ}] - [-75 \text{ kJ}] = -966 \text{ kJ} + 75 \text{ kJ} = \mathbf{-891 \text{ kJ}}$$

The combustion of 1 mol CH_4 releases 891 kJ.



APPLICATIONS 35

o6

Use standard enthalpies of formation from [Table R-11](#) to calculate $\Delta H_{\text{rxn}}^\circ$ for the following reaction.



APPLICATIONS 35

o6

Use standard enthalpies of formation from [Table R-11](#) to calculate $\Delta H_{\text{rxn}}^\circ$ for the following reaction.



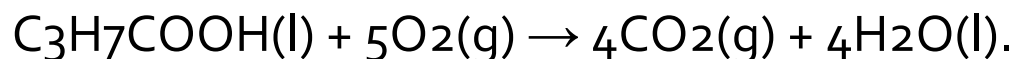
$$\Delta H_{\text{rxn}}^\circ = [4H_f^\circ(\text{NO}_2) + 6H_f^\circ(\text{H}_2\text{O})] - 4H_f^\circ(\text{NO}_3)$$

$$\begin{aligned}\Delta H_{\text{rxn}}^\circ &= [4(33.18 \text{ kJ}) + 6(-285.83 \text{ kJ})] \\ &\quad - 4(-46.11) \text{ kJ} \quad \Delta H_{\text{rxn}}^\circ = -1398 \text{ kJ}\end{aligned}$$

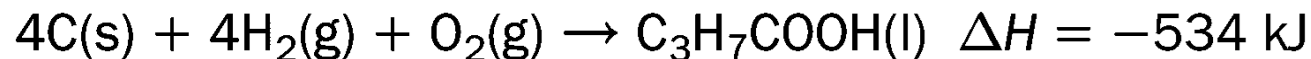
APPLICATIONS 36

07

Determine $\Delta H_{\text{comb}}^\circ$ for butanoic acid,



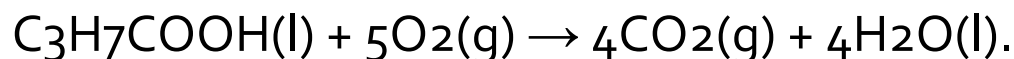
Use data in [Table R-11](#) and the following equation.



APPLICATIONS 36

07

Determine $\Delta H_{\text{comb}}^{\circ}$ for butanoic acid,



Use data in [Table R-11](#) and the following equation.



$$\Delta H_{\text{comb}}^{\circ} = [4H_f^{\circ}(\text{H}_2\text{O}) + 4H_f^{\circ}(\text{CO}_2)] \\ - H_f^{\circ}(\text{C}_3\text{H}_7\text{COOH})$$

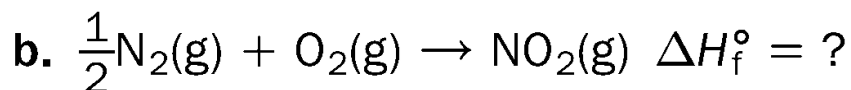
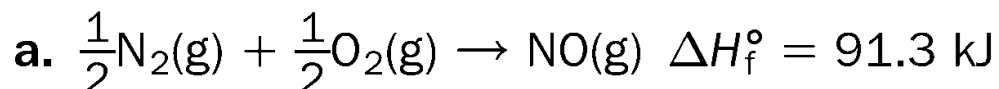
$$\Delta H_{\text{comb}}^{\circ} = [4(-286 \text{ kJ}) + 4(-394 \text{ kJ})] - (-534 \text{ kJ})$$

$$\Delta H_{\text{comb}}^{\circ} = -2186 \text{ kJ}$$

APPLICATIONS 37

o8

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:

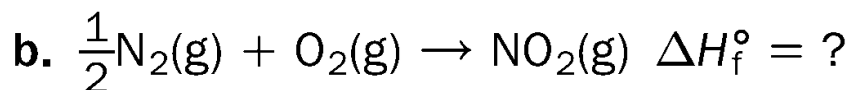
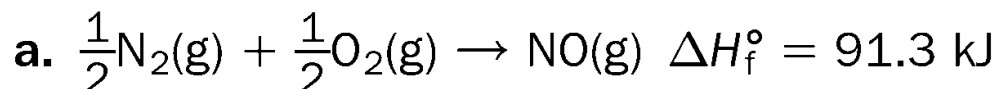


What is $\Delta H_{\text{f}}^{\circ}$ for Equation **b**?

APPLICATIONS 37

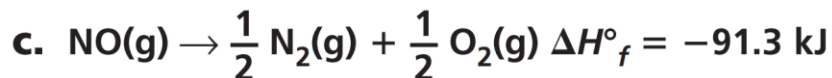
o8

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:

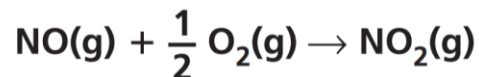


What is ΔH_f° for Equation b?

Reverse equation. a and change the sign of ΔH_f° to obtain equation c:



Add equations b and c:



$$\Delta H_{\text{rxn}}^{\circ} = -58.1 \text{ kJ} = \Delta H_f^{\circ}(\text{c}) + \Delta H_f^{\circ}(\text{b})$$

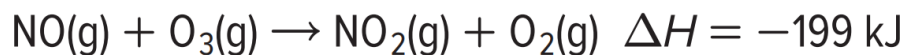
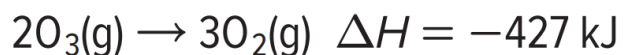
$$-58.1 \text{ kJ} = -91.3 \text{ kJ} + \Delta H_f^{\circ}(\text{b})$$

$$\Delta H_f^{\circ}(\text{b}) = -58.1 \text{ kJ} + 91.3 \text{ kJ} = 33.2 \text{ kJ}$$

REVIEW 42

09

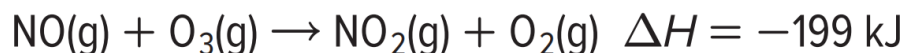
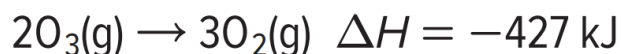
Use Hess's law to determine ΔH for the reaction $\text{NO(g)} + \text{O(g)} \rightarrow \text{NO}_2\text{(g)}$
 $\Delta H = ?$ given the following reactions.



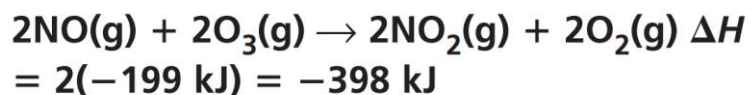
REVIEW 42

09

Use Hess's law to determine ΔH for the reaction $\text{NO(g)} + \text{O(g)} \rightarrow \text{NO}_2\text{(g)}$
 $\Delta H = ?$ given the following reactions.



Multiply c by 2:



Reverse b and change the sign of ΔH :



Reverse a and change the sign of ΔH :



Add the three equations and their ΔH values:



This is the equation and ΔH for 2 moles of NO reacting. Divide the equation and ΔH by 2:



How does ΔH for a thermochemical equation change when the amounts of all substances are tripled and the equation is reversed?

- a) ΔH is quadrupled, and its sign is changed.
- b) ΔH is tripled.
- c) ΔH is tripled, and its sign is changed.
- d) The sign of ΔH is changed.

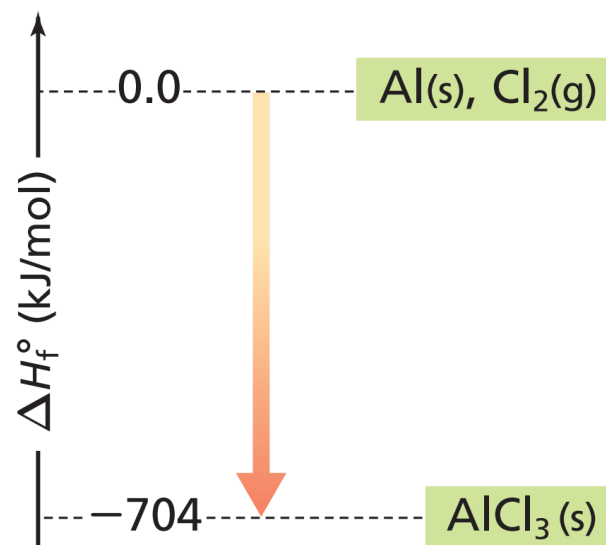
How does ΔH for a thermochemical equation change when the amounts of all substances are tripled and the equation is reversed?

- a) ΔH is quadrupled, and its sign is changed.
- b) ΔH is tripled.
- c) **ΔH is tripled, and its sign is changed.**
- d) The sign of ΔH is changed.

Mastering Concepts 91

11

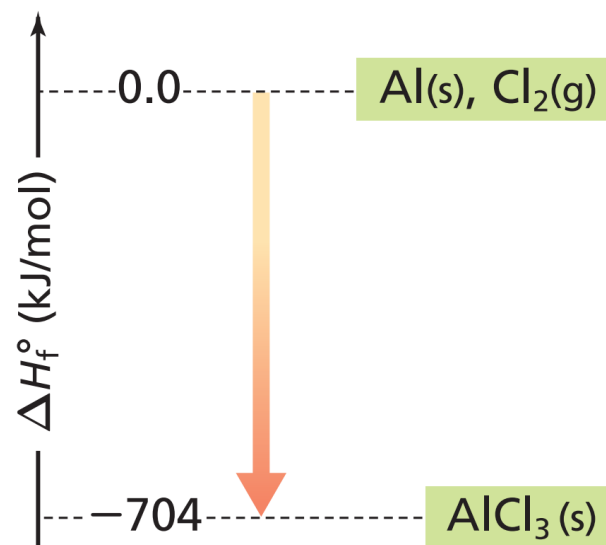
Use the figure to write the thermochemical equation for the formation of **1 mol** of aluminum chloride (a solid in its standard state) from its constituent elements in their standard states.



Mastering Concepts 91

11

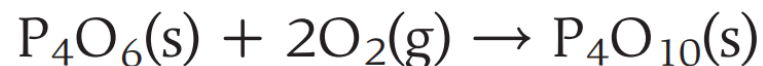
Use the figure to write the thermochemical equation for the formation of **1 mol** of aluminum chloride (a solid in its standard state) from its constituent elements in their standard states.



Mastering Problems 92

12

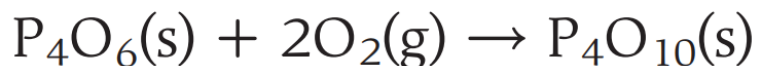
Use standard enthalpies of formation from [Table R-11](#) to calculate $\Delta H_{\text{rxn}}^\circ$ for the following reaction.



Mastering Problems 92

12

Use standard enthalpies of formation from [Table R-11](#) to calculate $\Delta H_{\text{rxn}}^\circ$ for the following reaction.



$$\Delta H_{\text{rxn}}^\circ = \Sigma \Delta H_f^\circ(\text{products}) - \Sigma \Delta H_f^\circ(\text{reactants})$$

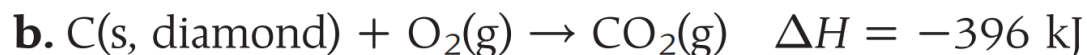
$$\begin{aligned}\Delta H_{\text{rxn}}^\circ &= [1(-2984.0 \text{ kJ})] - [1(-1640.1 \text{ kJ})] \\ &= -1343.9 \text{ kJ}\end{aligned}$$

Mastering Problems 93

13

Use Hess's law and the following thermochemical equations to produce the thermochemical equation for the reaction

$\text{C (s, diamond)} \rightarrow \text{C (s, graphite)}$. What is ΔH for the reaction?

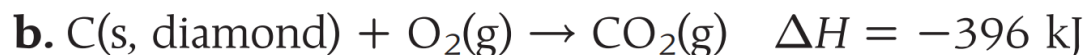


Mastering Problems 93

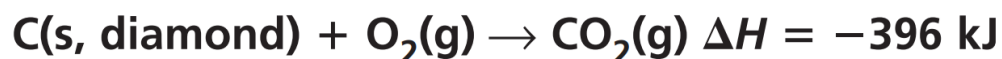
13

Use Hess's law and the following thermochemical equations to produce the thermochemical equation for the reaction

$\text{C (s, diamond)} \rightarrow \text{C (s, graphite)}$. What is ΔH for the reaction?



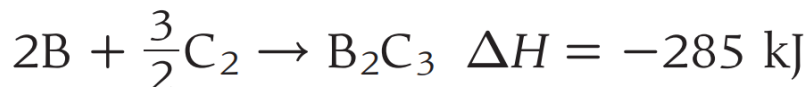
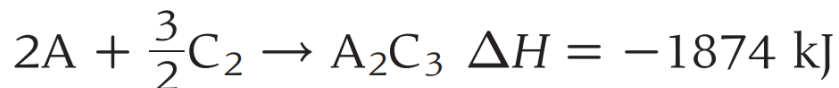
Reverse Equation a, and add to Equation b.



Mastering Problems 94

14

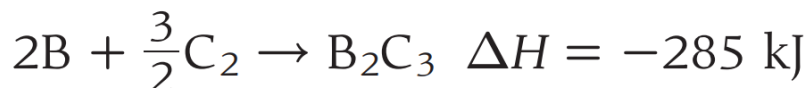
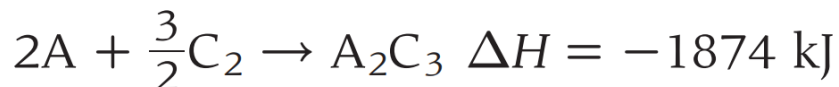
Use Hess's law and the changes in enthalpy for the following two generic reactions to calculate ΔH for the reaction $2A + B_2C_3 \rightarrow 2B + A_2C_3$.



Mastering Problems 94

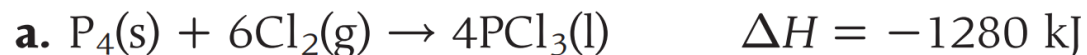
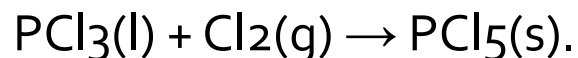
14

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

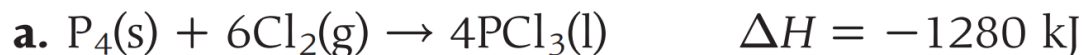
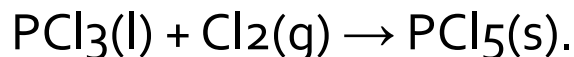


Reverse the second equation and change the sign of its ΔH value. Add the resulting equation to the first equation. Add the ΔH values. The resulting thermochemical equation is $2A + B_2C_3 \rightarrow 2B + A_2C_3 \quad \Delta H = -1589 \text{ kJ}$

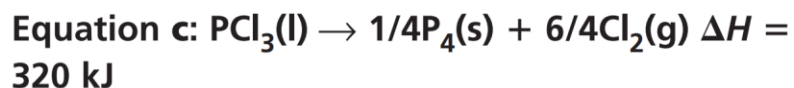
Demonstrate how thermochemical equations **a** and **b** can be used to determine the enthalpy change for the reaction



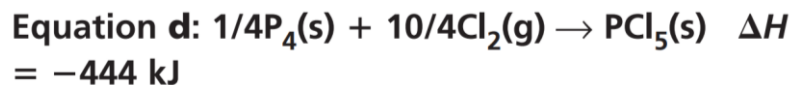
Demonstrate how thermochemical equations **a** and **b** can be used to determine the enthalpy change for the reaction



Reverse equation **a** and divide it by 4, yielding equation **c**.



Divide equation **b** by 4, yielding equation **d**.



Add equations **c** and **d** and their ΔH values.



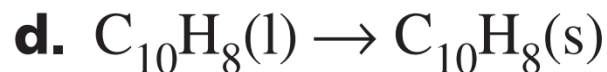
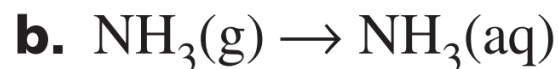
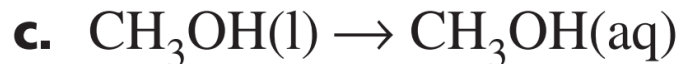
Reaction Spontaneity

Chapter 1 Lesson 5 – Revision Paper

APPLICATIONS 44

01

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



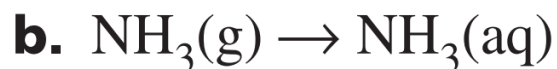
APPLICATIONS 44

01

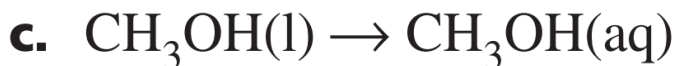
Predict the sign of ΔS_{system} for each of the following changes.



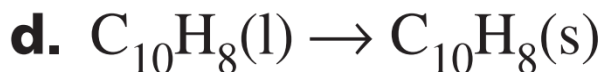
ΔS_{system} is negative because the system's entropy decreases. There are more gaseous reactant particles than product particles.



ΔS_{system} is negative because the system's entropy decreases. Aqueous particles have less freedom to move around.



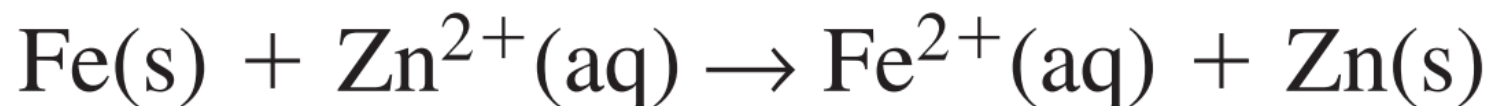
ΔS_{system} is positive because the system's entropy increases. Entropy increases when a solid or liquid dissolves to form a solution.



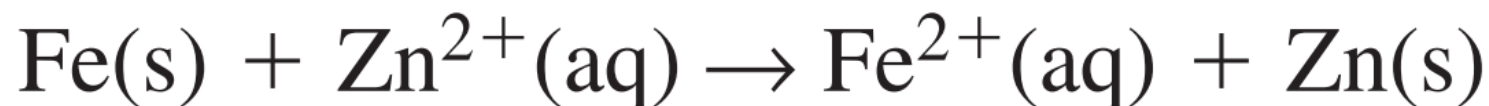
ΔS_{system} is negative because the system's entropy decreases. Solid particles have less freedom to move around than liquid particles.



Challenge Comment on the sign of ΔS_{system} for the following reaction.



Challenge Comment on the sign of ΔS_{system} for the following reaction.



The states of the two reactants are the same on both sides of the equation, so it is impossible from the equation alone to predict the sign of ΔS_{system} .

Example 7

03

For a process, $\Delta H_{\text{system}} = 145 \text{ kJ}$ and $\Delta S_{\text{system}} = 322 \text{ J/K}$. Is the process spontaneous at **382 K**?

Example 7

03

For a process, $\Delta H_{\text{system}} = 145 \text{ kJ}$ and $\Delta S_{\text{system}} = 322 \text{ J/K}$. Is the process spontaneous at **382 K**?

Convert ΔS_{system} to kJ/K

$$322 \cancel{\text{J}}/\text{K} \times \frac{1 \text{ kJ}}{1000 \cancel{\text{J}}} = 0.322 \text{ kJ/K}$$

Solve the free energy equation.

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

$$\Delta G_{\text{system}} = 145 \text{ kJ} - (382 \cancel{\text{K}})(0.322 \text{ kJ}/\cancel{\text{K}})$$

$$\Delta G_{\text{system}} = 145 \text{ kJ} - 123 \text{ kJ} = \mathbf{22 \text{ kJ}}$$

Because ΔG_{system} is positive, the reaction is nonspontaneous.

Convert ΔS_{system} to kJ/K.

State the Gibbs free energy equation.

Substitute $T = 382 \text{ K}$, $\Delta H_{\text{system}} = 145 \text{ kJ}$,
and $\Delta S_{\text{system}} = 0.322 \text{ kJ/K}$

Multiply and subtract numbers.

APPLICATIONS 46

04

Determine whether each of the following reactions is spontaneous.

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

APPLICATIONS 46

04

Determine whether each of the following reactions is spontaneous.

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

$$\Delta S_{\text{system}} = 138 \text{ J/K} = 0.138 \text{ kJ/K}$$

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

$$\Delta G_{\text{system}} = -75.9 \text{ kJ} - (273 \text{ K})(0.138 \text{ kJ/K})$$

$$\Delta G_{\text{system}} = -75.9 \text{ kJ} - 37.7 \text{ kJ} = -114 \text{ kJ}$$

spontaneous reaction



APPLICATIONS 46

04

Determine whether each of the following reactions is spontaneous.

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



APPLICATIONS 46

04

Determine whether each of the following reactions is spontaneous.

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

$$\Delta S_{\text{system}} = -55.2 \text{ J/K} = -0.0552 \text{ kJ/K}$$

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

$$\Delta G_{\text{system}} = -27.6 \text{ kJ} - (535 \text{ K})(-0.0552 \text{ kJ/K})$$

$$\Delta G_{\text{system}} = -27.6 \text{ kJ} + 29.5 \text{ kJ} = 1.9 \text{ kJ}$$

nonspontaneous reaction



APPLICATIONS 46

04

Determine whether each of the following reactions is spontaneous.

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

APPLICATIONS 46

04

Determine whether each of the following reactions is spontaneous.

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

$$\Delta S_{\text{system}} = -55.2 \text{ J/K} = -0.0552 \text{ kJ/K}$$

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

$$\Delta G_{\text{system}} = 365 \text{ kJ} - (388 \text{ K})(-0.0552 \text{ kJ/K})$$

$$\Delta G_{\text{system}} = 365 \text{ kJ} + 21.4 \text{ kJ} = 386 \text{ kJ}$$

nonspontaneous reaction



APPLICATIONS 46

04

Determine whether each of the following reactions is spontaneous.

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

APPLICATIONS 46

04

Determine whether each of the following reactions is spontaneous.

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

$$\Delta S_{\text{system}} = 55.7 \text{ J/K} = 0.0557 \text{ kJ}$$

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

$$\Delta G_{\text{system}} = 452 \text{ kJ} - (165 \text{ K})(0.0557 \text{ kJ/K})$$

$$\Delta G_{\text{system}} = 452 \text{ kJ} - 9.19 \text{ kJ} = 443 \text{ kJ}$$

nonspontaneous reaction



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

APPLICATIONS 47

05

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

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

For the reaction to be spontaneous:

$$\Delta G_{\text{system}} < 0: \Delta H_{\text{system}} - T\Delta S_{\text{system}} < 0$$

$$T < \frac{\Delta H_{\text{system}}}{\Delta S_{\text{system}}}$$

$$T < \frac{-144 \text{ kJ}}{(-36.8 \text{ J/K})(1 \text{ kJ}/1000 \text{ J})}$$

$$T < 3910 \text{ K}$$

REVIEW 49

o6

Describe how a system's entropy changes if the system becomes more disordered during a process.

- a) The system's entropy does not change.
- b) The system's entropy decreases.
- c) The system's entropy increases.
- d) The system's entropy is negative.

REVIEW 49

o6

Describe how a system's entropy changes if the system becomes more disordered during a process.

- a) The system's entropy does not change.
- b) The system's entropy decreases.
- c) **The system's entropy increases.**
- d) The system's entropy is negative.

REVIEW 50

07

Does the entropy of a system increase or decrease when you dissolve a cube of sugar in a cup of tea? Define the system and explain your answer.

Does the entropy of a system increase or decrease when you dissolve a cube of sugar in a cup of tea? Define the system and explain your answer.

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.

REVIEW 51

o8

Determine whether the system $\Delta H_{\text{system}} = -20.5 \text{ kJ}$, $T = 298 \text{ K}$, and $\Delta S_{\text{system}} = -35.0 \text{ J/K}$ is spontaneous or nonspontaneous.

REVIEW 51

o8

Determine whether the system $\Delta H_{\text{system}} = -20.5 \text{ kJ}$, $T = 298 \text{ K}$, and $\Delta S_{\text{system}} = -35.0 \text{ J/K}$ is spontaneous or nonspontaneous.

$$\Delta S_{\text{system}} = -35.0 \text{ J/K} = -0.0350 \text{ kJ/K}$$

$$\Delta G_{\text{system}} = -20.5 \text{ kJ} - (298 \text{ K})(-0.0350 \text{ kJ/K}) = -10.1 \text{ kJ}$$

The system is spontaneous.



Under what conditions is an endothermic chemical reaction in which the entropy of the system increases likely to be spontaneous?

- a) At higher temperatures
- b) At lower temperatures
- c) When the sign of the free energy is positive
- d) When the sign of the entropy change is positive

Under what conditions is an endothermic chemical reaction in which the entropy of the system increases likely to be spontaneous?

- a) **At higher temperatures**
- b) At lower temperatures
- c) When the sign of the free energy is positive
- d) When the sign of the entropy change is positive

Predict how the entropy of the system changes for the reaction



- a) Doesn't change
- b) Increases
- c) Decreases
- d) Can't be determined

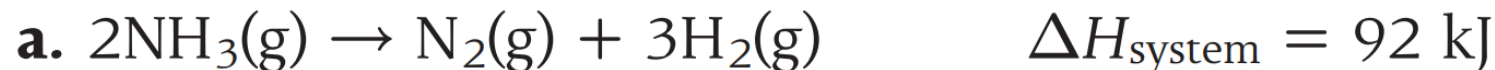
Predict how the entropy of the system changes for the reaction



- a) Doesn't change
- b) Increases**
- c) Decreases
- d) Can't be determined

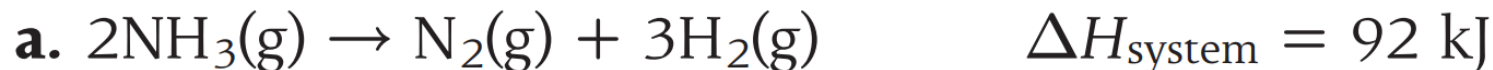
Because a gaseous product is formed, it's likely that the system's entropy increases.

Which of these reactions would you expect to be spontaneous at relatively high temperatures?



- a) a
- b) a and b
- c) c
- d) a and c

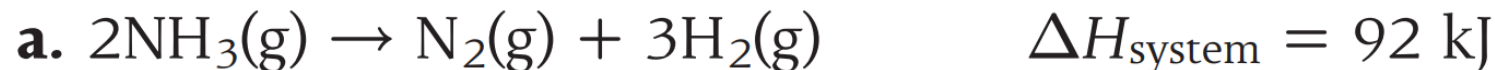
Which of these reactions would you expect to be spontaneous at relatively high temperatures?



- a) a
- b) a and b
- c) c
- d) a and c

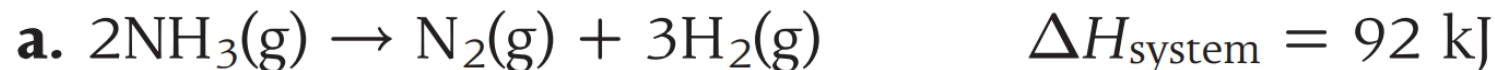
For a spontaneous reaction, ΔG_{system} must be negative as calculated in the expression $\Delta G_{\text{system}}^{\circ} = \Delta H_{\text{system}}^{\circ} - T\Delta S_{\text{system}}^{\circ}$. Reactions a and c both have a positive $\Delta H_{\text{system}}^{\circ}$. However, both reactions also have more moles of gaseous products than gaseous reactants, which suggests that entropy increases as products form. So, higher temperatures will tend to make these reactions spontaneous (a negative value for $\Delta G_{\text{system}}^{\circ}$).

Which of these reactions would you expect to be spontaneous at relatively low temperatures?



- a) c
- b) b
- c) a
- d) b and c

Which of these reactions would you expect to be spontaneous at relatively low temperatures?



a) c

b) b

c) a

d) b and c

reaction **b** has fewer moles of gas on the products side, which means entropy decreases as products form. But because ΔH_{system} is negative for this reaction, it will tend to be spontaneous at lower temperatures.

Mastering Concepts 98

12

Explain how an exothermic reaction changes the entropy of the surroundings. Does the enthalpy change for such a reaction increase or decrease ΔG_{system} ? Explain.

Mastering Concepts 98

12

Explain how an exothermic reaction changes the entropy of the surroundings. Does the enthalpy change for such a reaction increase or decrease ΔG_{system} ? Explain.

The heat released by an exothermic reaction increases the entropy of the surroundings. Such a reaction decreases ΔG_{system} because ΔH_{system} is negative in the equation $\Delta G^{\circ}_{\text{system}} = \Delta H^{\circ}_{\text{system}} - T\Delta S^{\circ}_{\text{system}}$.

Mastering Problems 99

13

Calculate ΔG_{system} for each process, and state whether the process is spontaneous or nonspontaneous.

a. $\Delta H_{\text{system}} = 145 \text{ kJ}$, $T = 293 \text{ K}$, $\Delta S_{\text{system}} = 195 \text{ J/K}$

b. $\Delta H_{\text{system}} = -232 \text{ kJ}$, $T = 273 \text{ K}$, $\Delta S_{\text{system}} = 138 \text{ J/K}$

c. $\Delta H_{\text{system}} = -15.9 \text{ kJ}$, $T = 373 \text{ K}$, $\Delta S_{\text{system}} = -268 \text{ J/K}$

Mastering Problems 99

13

Calculate ΔG_{system} for each process, and state whether the process is spontaneous or nonspontaneous.

a. $\Delta H_{\text{system}} = 145 \text{ kJ}$, $T = 293 \text{ K}$, $\Delta S_{\text{system}} = 195 \text{ J/K}$

b. $\Delta H_{\text{system}} = -232 \text{ kJ}$, $T = 273 \text{ K}$, $\Delta S_{\text{system}} = 138 \text{ J/K}$

c. $\Delta H_{\text{system}} = -15.9 \text{ kJ}$, $T = 373 \text{ K}$, $\Delta S_{\text{system}} = -268 \text{ J/K}$

a. $\Delta H_{\text{system}} = 145 \text{ kJ}$, $T = 293 \text{ K}$, $\Delta S_{\text{system}} = 195 \text{ J/K}$

$$\Delta S_{\text{system}} = 195 \text{ J/K} = 0.195 \text{ kJ/K}$$

$$\Delta G^{\circ}_{\text{system}} = \Delta H^{\circ}_{\text{system}} - T\Delta S^{\circ}_{\text{system}}$$

$$\Delta G^{\circ}_{\text{system}} = 145 \text{ kJ} - (293\text{K})(0.195 \text{ kJ/K}) \\ = 87.9 \text{ kJ}$$

nonspontaneous

b. $\Delta H_{\text{system}} = -232 \text{ kJ}$, $T = 273 \text{ K}$, $\Delta S_{\text{system}} = 138 \text{ J/K}$

$$\Delta S_{\text{system}} = 0.138 \text{ kJ/K}$$

$$\Delta G^{\circ}_{\text{system}} = -232 \text{ kJ} - (273\text{K})(0.138 \text{ kJ/K}) \\ = -270 \text{ kJ}$$

spontaneous

c. $\Delta H_{\text{system}} = -15.9 \text{ kJ}$, $T = 373 \text{ K}$, $\Delta S_{\text{system}} = -268 \text{ J/K}$

$$\Delta S_{\text{system}} = -268 \text{ J/K} = -0.268 \text{ kJ/K}$$

$$\Delta G^{\circ}_{\text{system}} = -15.9 \text{ kJ} - (373\text{K})(-0.268 \text{ kJ/K}) \\ = 84.1 \text{ kJ}$$

nonspontaneous

Mastering Problems 100

14

Calculate the temperature at which $\Delta G_{\text{system}} = 0$ if $\Delta H_{\text{system}} = 4.88 \text{ kJ}$ and $\Delta S_{\text{system}} = 55.2 \text{ J/K}$.

Mastering Problems 100

14

Calculate the temperature at which $\Delta G_{\text{system}} = 0$ if $\Delta H_{\text{system}} = 4.88 \text{ kJ}$ and $\Delta S_{\text{system}} = 55.2 \text{ J/K}$.

$$\Delta S_{\text{system}} = 55.2 \text{ J/K} = 0.0552 \text{ kJ/K}$$

$$\Delta G = \Delta H - T\Delta S$$

$$0 = 4.88 \text{ kJ} - T(0.0552 \text{ kJ/K})$$

$$T = \frac{4.88 \text{ kJ}}{0.0552 \text{ kJ/K}} = 88.4 \text{ K}$$



For the change $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$, $\Delta G_{\text{system}}^\circ$ is **8.557 kJ** and $\Delta H_{\text{system}}^\circ$ is **44.01 kJ**. What is $\Delta S_{\text{system}}^\circ$ for the change?

For the change $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$, $\Delta G_{\text{system}}^\circ$ is **8.557 kJ** and $\Delta H_{\text{system}}^\circ$ is **44.01 kJ**. What is $\Delta S_{\text{system}}^\circ$ for the change?

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$8.557 \text{ kJ} = 44.01 \text{ kJ} - (298 \text{ K}) \Delta S^\circ$$

$$\Delta S^\circ = -\frac{8.557 \text{ kJ} - 44.01 \text{ kJ}}{298 \text{ K}} = 0.119 \text{ kJ/K}$$



Is the reaction to convert copper(II) sulfide to copper(II) sulfate spontaneous under standard conditions? $\text{CuS(s)} + 2\text{O}_2\text{(g)} \rightarrow \text{CuSO}_4\text{(s)}$. $\Delta H_{\text{rxn}}^\circ = -718.3 \text{ kJ}$, and $\Delta S_{\text{rxn}}^\circ = -368 \text{ J/K}$. Explain.

Is the reaction to convert copper(II) sulfide to copper(II) sulfate spontaneous under standard conditions? $\text{CuS(s)} + 2\text{O}_2\text{(g)} \rightarrow \text{CuSO}_4\text{(s)}$. $\Delta H_{\text{rxn}}^\circ = -718.3 \text{ kJ}$, and $\Delta S_{\text{rxn}}^\circ = -368 \text{ J/K}$. Explain.

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$\Delta G^\circ = -718.3 \text{ kJ} - (298 \text{ K})(-0.368 \text{ kJ/K}); \Delta G^\circ = -609 \text{ kJ}$$

Yes. The reaction is spontaneous under standard conditions because $\Delta G_{\text{rxn}}^\circ = -609 \text{ kJ}$, and a negative $\Delta G_{\text{rxn}}^\circ$ indicates spontaneity.

Mastering Problems 103

17

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

Mastering Problems 103

17

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

$$\Delta G^{\circ}_{\text{system}} = \Delta H^{\circ}_{\text{system}} - T\Delta S^{\circ}_{\text{system}}$$

$$-34.7 \text{ kJ} = -28.8 \text{ kJ} - T(0.0222 \text{ kJ/K})$$

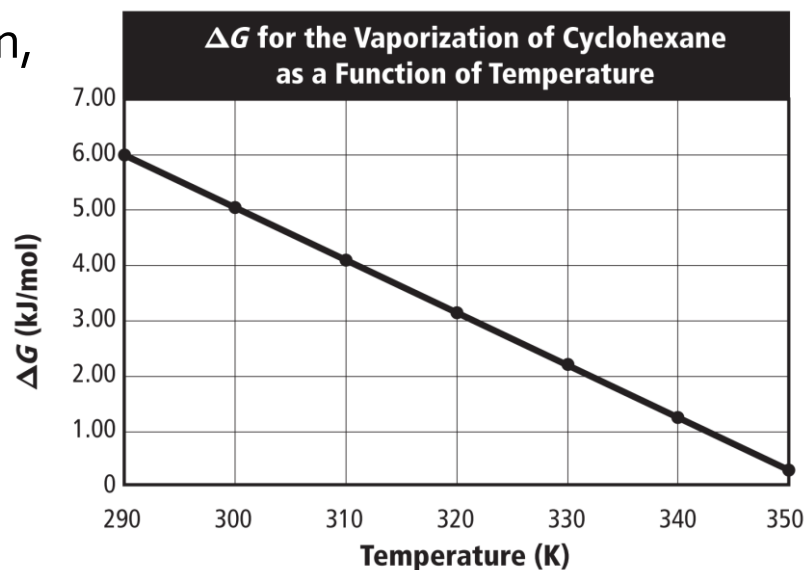
$$T = 266 \text{ K}$$

Multiple Choice 1

18

In the range of temperatures shown, the vaporization of cyclohexane

- a) does not occur at all.
- b) will occur spontaneously.
- c) is not spontaneous.
- d) occurs only at high temperatures.

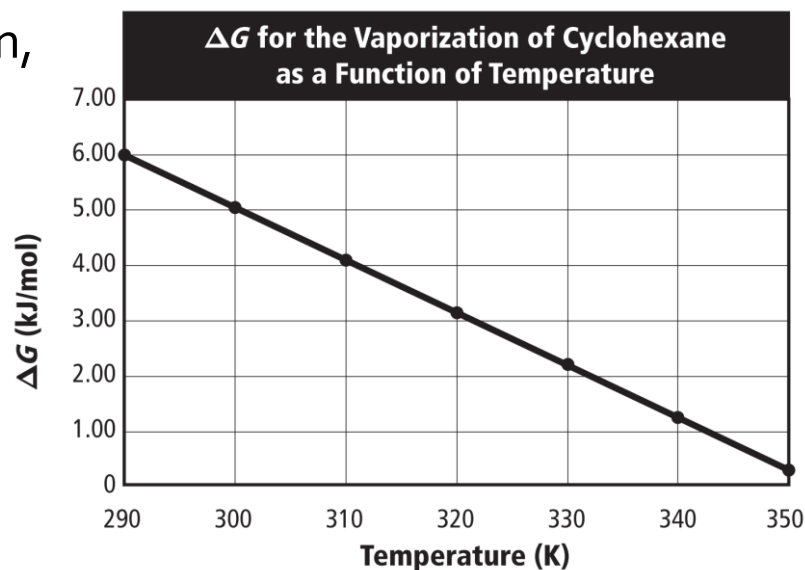


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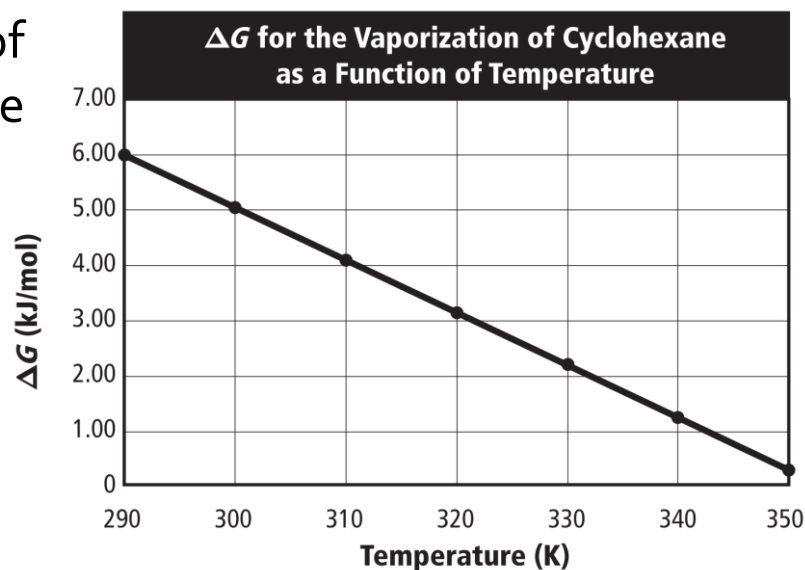


Multiple Choice 2

19

What is the standard free energy of vaporization, $\Delta G_{\text{vap}}^{\circ}$, of cyclohexane at 300 K?

- a) 5.00 kJ/mol
- b) 3.00 kJ/mol
- c) 4.00 kJ/mol
- d) 2.00 kJ/mol

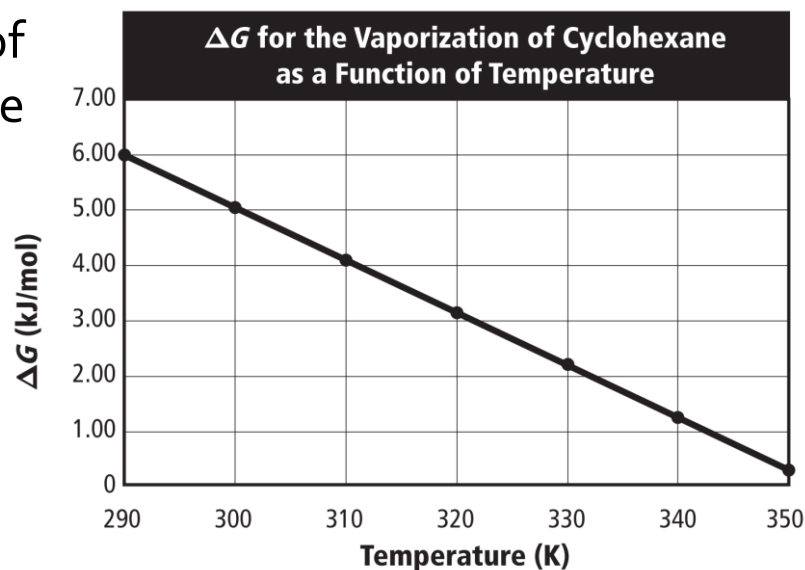


Multiple Choice 2

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- c) 4.00 kJ/mol
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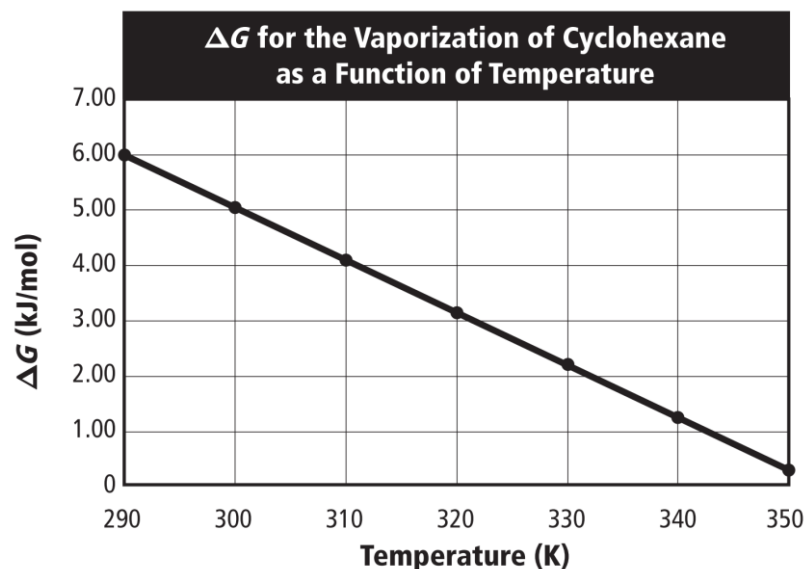


Multiple Choice 3

20

When $\Delta G_{\text{vap}}^{\circ}$ is plotted as a function of temperature, the slope of the line equals $\Delta S_{\text{vap}}^{\circ}$ and the y-intercept of the line equals $\Delta H_{\text{vap}}^{\circ}$. What is the approximate standard entropy of the vaporization of cyclohexane?

- a) $-50.0 \text{ J/mol}\cdot\text{K}$
- b) $-5.00 \text{ J/mol}\cdot\text{K}$
- c) $-10.0 \text{ J/mol}\cdot\text{K}$
- d) $-100 \text{ J/mol}\cdot\text{K}$

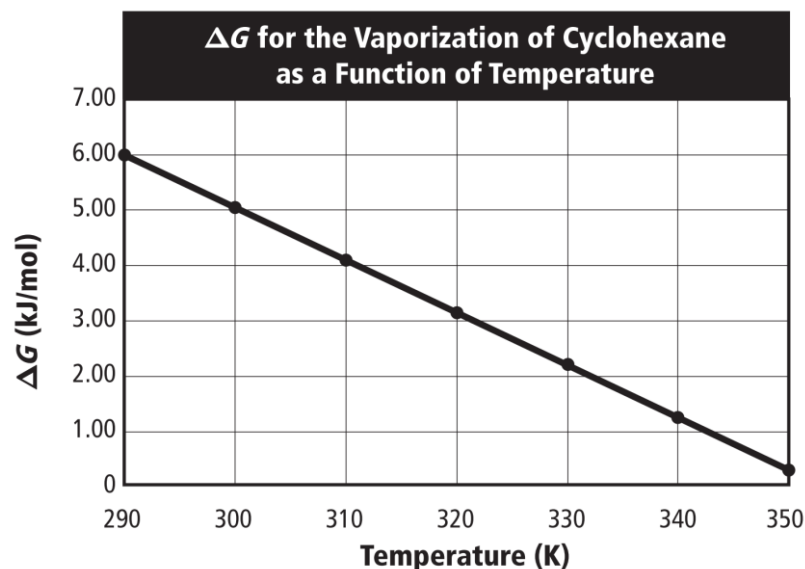


Multiple Choice 3

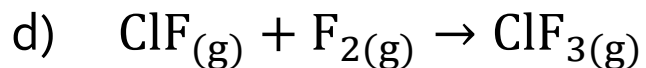
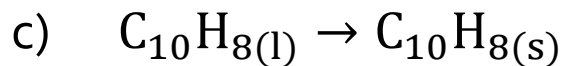
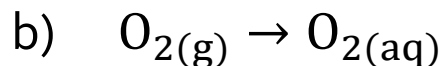
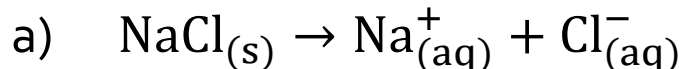
20

When $\Delta G_{\text{vap}}^{\circ}$ is plotted as a function of temperature, the slope of the line equals $\Delta S_{\text{vap}}^{\circ}$ and the y-intercept of the line equals $\Delta H_{\text{vap}}^{\circ}$. What is the approximate standard entropy of the vaporization of cyclohexane?

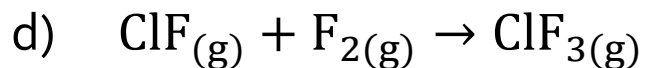
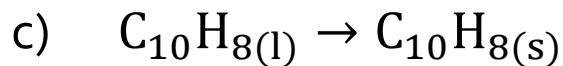
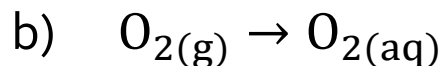
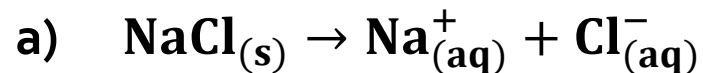
- a) $-50.0 \text{ J/mol}\cdot\text{K}$
- b) $-5.00 \text{ J/mol}\cdot\text{K}$
- c) $-10.0 \text{ J/mol}\cdot\text{K}$
- d) **$-100 \text{ J/mol}\cdot\text{K}$**



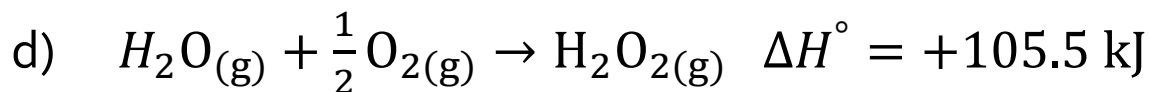
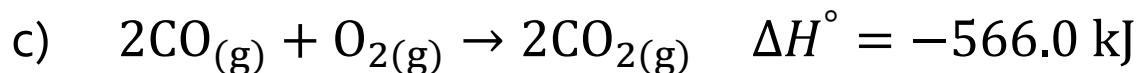
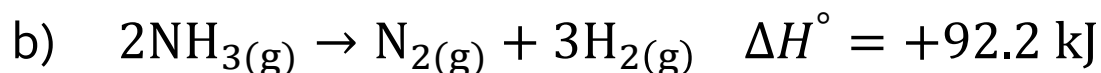
Which of the following processes the $\Delta S_{\text{system}}^{\circ}$ sign is **positive**?



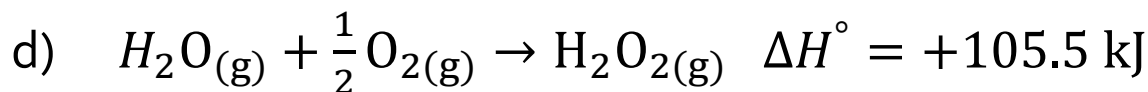
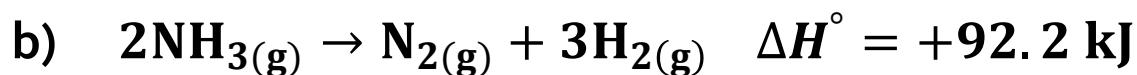
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Which of the following reactions would you expect to be spontaneous at high temperature?



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EOT1 20-21 Exam Question

23

What is the ΔG_{system} value of the following process?

$$\Delta S_{\text{system}}^{\circ} = 195 \text{ J/K}, \quad \Delta H_{\text{system}}^{\circ} = 145 \text{ kJ} \quad \text{and} \quad T = 293 \text{ K}$$

- a) $-5.7 \times 10^4 \text{ kJ}$
- b) $-4.25 \times 10^4 \text{ kJ}$
- c) 87.9 kJ
- d) 152.5 kJ



EOT1 20-21 Exam Question

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Chapter 1 – Revision Paper

Resources

- *Energy and Chemical Change, Chapter 15, from Glencoe Chemistry: Matter and Change ©2017*

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