**Name ______________________________** **Honors Chemistry** **___/___/___**

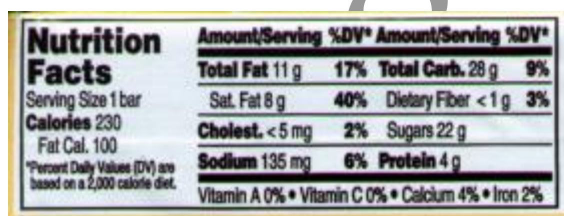
## Energy

**Energy** is the ability to do work or produce heat. Energy can be classified as either **kinetic energy** (energy of motion) or **potential energy** (energy due to position or composition).

One step in understanding energy is realizing the difference between **heat** and **temperature**. Temperature is a measure of the average kinetic energy of the particles in a substance. Heat is the flow of energy due to a temperature difference. Heat travels from a higher temperature object to a lower temperature object until both objects reach the same temperature. The **law of conservation of energy** states that energy cannot be created or destroyed. So in all heat transfers one object is **endothermic** (absorb heat) and the other is **exothermic** (release heat).

In chemistry we measure heat using two different units: calories and joules. A **calorie** is defined as the amount of energy needed to raise the temperature of one gram of water one degree Celsius. The "calorie" you are most familiar with is used to measure the energy content of food. This "calorie" is actually a kilocalorie (1000 calories) and is written with a capital C to distinguish it from the calorie used in chemistry.

The Calories listed for this candy bar are actually kilocalories.



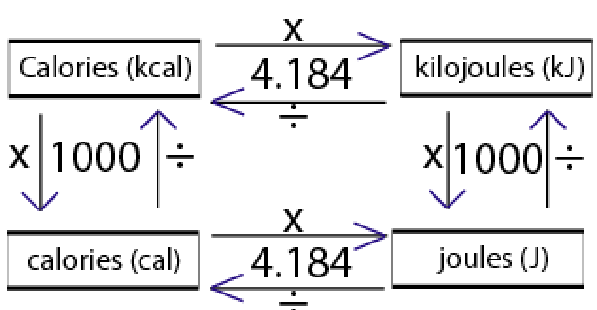

The other unit used to measure heat is the **joule**. The joule is the SI (System International) unit.


In this juice label from Amsterdam you can see that energy ("energie") is given in kilojoules and kilocalories.

To convert between calories and joules keep in mind:

1 calorie (cal) = 4.184 joules (J)
or
1 Calorie (kcal) = 1000 calories (cal) = 4184 joules (J) = 4.184 kilojoules (kJ)
or use this chart:

(you will need to **memorize** one of these conversion methods)



## Specific Heat Capacity

**Specific heat capacity (C)** is defined as the amount of energy required to raise the temperature of one gram of a substance one degree Celsius. The unit for specific heat capacity is J/g°C. Specific heat capacity is sometimes referred to as just specific heat. Specific heat capacity values vary for different substances. It is a substance's specific heat that makes some substances feel cooler than others, even though both are at the same temperature. This phenomenon is due to the fact that some substances require little heat in order for their temperature to rise 1°C. Thus, substances with low specific heat values are said to be good conductors of heat. Most metals are considered to be good conductors because of their low C values. Substances with high C values are called insulators because they require large amounts of heat in order to warm up just a little. It is because of this, large bodies of water take a long time to warm in the summer and a long time to freeze in the winter. The chart below lists the specific heat capacities for some various substances.

| Substance | Specific Heat Capacity (J/g°C) |
|---|---|
| Water (l) | 4.184 |
| Water (s) | 2.09 |
| Water (g) | 2.01 |
| Aluminum (s) | 0.89 |
| Iron (s) | 0.45 |
| Mercury (l) | 0.14 |
| Carbon (s) | 0.71 |
| Silver (s) | 0.24 |
| Gold (s) | 0.13 |
| Copper (s) | 0.385 |

The specific heat capacity of a substance is needed when determining how much energy is absorbed or released when heat flows from one substance to another. The formula used is:

**$q = mC\Delta T$**

The $\Delta$ is the delta symbol and means "change in".

**q** – heat, measured in joules

**m** – the mass of a substance, measured in grams

**C** – the specific heat capacity of the substance that is being heated or cooled, measured in J/g°C

**$\Delta$T** – the change in temperature (final temperature – initial temperature), measured in °C

Since you will either be given the initial and final temperatures or asked to find one of them, it may be better to express the formula as:

**$q = \text{mass} \times C \times (T_{final} - T_{initial})$**

If q is **positive**, it is an **endothermic change**, if q is **negative**, it is an **exothermic change**.

**Example:** How much heat energy is released to your body when a cup of hot tea containing 200.0 grams of water is cooled from 65°C to body temperature, 37°C?

**$q = \text{mass} \times C \times (T_{final} - T_{initial})$**

**q = (200.0g)(4.184 J/g °C)(37°C-65°C)**

**q = (200.0g)(4.184 J/g°C)(-28°C)**

**q = -23000 J = -23 kJ = -5.6 kcal**

**exothermic**

**Which C value of water should you use?**
Water boils at 100°C and freezes at 0°C. If a temperature range is below 0°C, use the C value for solid water. If the temperature range is above 100°C use the C value for water vapor (g). If the temperature range is between 0°C and 100°C, use the C value for liquid water.

## Phase Changes

Changes of state are physical changes where heat is absorbed or released. If a substance absorbs heat, the process is endothermic and it will change from a lower energy state to a higher energy state. If a substance releases heat, the process is exothermic and it will change from a higher energy state to a lower energy state. The lowest energy state is solid and the highest energy state is gas. The liquid state is intermediate between solid and gas.

| Name | Change | Endo/Exo |
|---|---|---|
| Melting | Solid to Liquid | Endo |
| Freezing | Liquid to Solid | Exo |
| Sublimation | Solid to Gas | Endo |
| Deposition | Gas to Solid | Exo |
| Vaporization | Liquid to Gas | Endo |
| Condensation | Gas to Liquid | Exo |

As a substance melts, freezes, boils or condenses, heat is either absorbed or released. In molecular (covalent) compounds bonds between molecules are broken or formed; in ionic compounds bonds between ions are broken or formed. As the bonds between molecules, ions or atoms break, the particles move further and further apart; when bonds form, particles move closer and closer together.

During physical changes of state, **there is no change in temperature**. Thus all energy being added is used to break bonds and any energy released is from the formation of bonds. For each change of state there is a term used to describe the process and an energy value associated with it.

The chart below summarizes these processes:

| Process | Name | Symbol | Endo/Exo | Bonds | Value for Water |
|---|---|---|---|---|---|
| Melting (s→l) | Heat of fusion | $H_{fus}$ | Endo | Broken | 335 J/g or 80. cal/g |
| Freezing (l→s) | Heat of solidification | $H_{sol}$ | Exo | Formed | -335 J/g or -80. cal/g |
| Vaporization (l→g) | Heat of vaporization | $H_{vap}$ | Endo | Broken | 2259.0 J/g or 540. cal/g |
| Condensation (g→l) | Heat of condensation | $H_{con}$ | Exo | Formed | -2259.0 J/g or -540. cal/g |

*Each of the values listed above represent the amount of energy needed or released to (melt/freeze/vaporize/condense) **one gram** of water.

Because the temperature is constant, the formula: **$q = mC\Delta T$** would not be accurate. Without a temperature change, $\Delta T = 0$ and our calculation of the heat absorbed or released would be zero. Since we know heat is being added or released a different formula must be used. At the melting/freezing point the formula to be used is: **$q = H_x$ x mass (or moles)**.

**Example:** Calculate the amount of energy needed to melt 5.00 grams of ice at 0°C.
**$q = H_{fus}$ x mass**
**q = 335.0 x 5.00**
**q = 1680 joules ÷ 1000 = 1.68 kJ ÷ 4.184 = 0.402 kcal**

The reverse process(freezing) is an exothermic process. Thus, instead of $H_{fus}$, $H_{solid}$ (heat of solidification) is used. It has the same numeric value as a substance's $H_{fus}$ but it is negative ($H_{solid}$ ($H_2O$) = -335.0 J/g). If the reverse process (freezing) were to occur above, then q = -1680 J.

**Example:** Calculate the amount of energy needed to change 30.00 grams of water at 100°C to steam at 100°C.
**$q = H_{vap}$ x mass**
**q = 2259.0 x 30.00**
**q = 67770 joules ÷ 1000 = 67.77 kJ ÷ 4.184 = 16.20 kcal**

The reverse process(condensing) is an exothermic process. Thus, instead of $H_{vap}$, $H_{cond}$ (heat of condensation) is used. It has the same numeric value as a substance's $H_{vap}$ but it is negative ($H_{cond}$ ($H_2O$) = -2259.0 J/g). If the reverse process (condensing) were to occur above, then q = -67770 J.

**Note:** The $H_{fus}$, $H_{cond}$, $H_{vap}$ & $H_{solid}$ values given above are in joules/gram. On occasion, however, you will be given values in joules/mole or kilojoules/mole. If this is the case, use the formulas: **$q = \Delta H_{fus}$ x moles** or **$q = \Delta H_{vap}$ x moles. Be sure to convert mass → moles if necessary.** See the example below.

**Example:** Calculate the kilojoules of energy needed to melt 10.0 grams of $NH_3$. $H_{fus}$ = 5.65 kJ/mole

First: Since the $H_{fus}$ value is given in kJ/mole, I will use the formula **$q = \Delta H_{fus}$ x moles.**
Second: Convert mass to moles by dividing the given value by the gram formula mass of the substance.
10.0 ÷ 17.04 = 0.587 moles
Third: Solve it.
**$q = \Delta H_{fus}$ x moles**
**q = 5.65 x 0.587**
**q = 3.32 kJ ÷ 4.184 = 0.793 kcal**

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**Energy – Answers**

**Homework:**
**Part 1: Solve each specific heat capacity problem in kilojoules and Calories (kilocalories) or degrees Celsius. Indicate if an endothermic or exothermic change takes place.**

1. If you fill a bathtub with 200. kg (convert kilograms to grams!) of water at 44°C, how much heat energy is lost as the water cools to a temperature of 21°C?

q = x
$T_{initial}$ = 44°C
$T_{final}$ = 21°C
Mass = 200 kilograms x 1000 = 200000 grams
C = 4.184 J/g°C

**q = mass x C x ($T_{final}$ – $T_{initial}$)**
**q = (200000) x (4.184) x (21 – 44)**
**q = -19246400 J**
**q = -19000 kJ**
**q = -4600 Calories**
**exothermic**

2. On a cold winter day with a temperature of 4°C, you pick up a penny from the ground and put it in your pocket. If the penny has a mass of 1.85 grams, how much heat energy must be transferred to the coin to warm it to your body temperature, 37°C?

q = x
$T_{initial}$ = 4°C
$T_{final}$ = 37°C
Mass = 1.85 grams
C = 0.385 J/g°C

**q = mass x C x ($T_{final}$ – $T_{initial}$)**
**q = (1.85) x (0.385) x (37 – 4)**
**q = 23.5 J**
**q = 0.024 kJ**
**q = 0.0056 Calories**
**endothermic**

3. Predict the final temperature of 2.50 kg (covert kilograms to grams) of water in a calorimeter if the water is at 25.0°C before 0.5 oz. of noodles, which contain 54 Calories (convert to joules), are burned?

q = 54 Calories x 1000 = 54000 calories x 4.184 = 225936 joules
$T_{initial}$ = 25.0 °C
$T_{final}$ = x
Mass = 2.50 kilograms x 1000 = 2500 grams
C = 4.184 J/g°C

**q = mass x C x ($T_{final}$ – $T_{initial}$)**
**225936 = (2500) x (4.184) x ($T_f$ – 25)**
**225936 = 10460 x ($T_f$ – 25)**
**21.6 = x – 25**
**x = 47° C**
**endothermic**

4. How much heat is lost by 4.0 L (convert L to mL & convert mL to grams; density of water 1.0 g/mL) of water that is cooled from 87°C to 21°C?

q = x
$T_{initial}$ = 87 °C
$T_{final}$ = 21°C
Mass = 4.0 L x 1000 = 4000 mL = 4000 grams
C = 4.184 J/g°C

**q = mass x C x ($T_{final}$ – $T_{initial}$)**
**q = (4000) x (4.184) x (21 – 87)**
**q = -1104576 J**
**q = -1.1 x $10^3$ kJ**
**q = -260 Calories**
**exothermic**

5. If 980 kJ of energy are added to 6.2 kg of water at 18°C, what is the final temperature of the water?

q = 980 kJ = 980000 joules
$T_{initial}$ = 18°C
$T_{final}$ = x
Mass = 6.2 kg = 6200 grams
C = 4.184 J/g°C

**q = mass x C x ($T_{final}$ – $T_{initial}$)**
**980000 = 6200 x 4.184 x ($T_f$ – 18)**
**980000 = 25940.8 x ($T_f$ – 18)**
**37.8 = x – 18**
**x = 56 °C**
**endothermic**

6. How much heat energy is needed to raise the temperature of a 425.0 gram aluminum baking sheet from room temperature, 25°C, to a baking temperature of 200.°C?

q = x
$T_{initial}$ = 25 °C
$T_{final}$ = 200.°C
Mass = 425.0 grams
C = 0.89 J/g°C

**q = mass x C x ($T_{final}$ – $T_{initial}$)**
**q = (425.0) x (0.89) x (200. – 25)**
**q = 66193.75 J**
**q = 66 kJ**
**q = 16 Calories**
**endothermic**

7. Calculate the amount of energy needed to raise the temperature of 100. grams of gold from 30.°C to 40.°C.

q = x
$T_{initial}$ = 30. °C
$T_{final}$ = 40.°C
Mass = 100 grams
C = 0.13 J/g°C

**q = mass x C x ($T_{final}$ – $T_{initial}$)**
**q = (100) x (0.13) x (40 – 30)**
**q = 130 J**
**q = 0.130 kJ**
**q = 0.031 Calories**
**endothermic**

8. Calculate the amount of energy released as 225 grams of mercury cools from 100.°C to 27°C.

q = x
$T_{initial}$ = 100 °C
$T_{final}$ = 27°C
Mass = 225 grams
C = 0.14 J/g°C

**q = mass x C x ($T_{final}$ – $T_{initial}$)**
**q = (225) x (0.14) x (27 – 100)**
**q = -2299.5 J**
**q = -2.3 kJ**
**q = -0.55 Calories**
**exothermic**

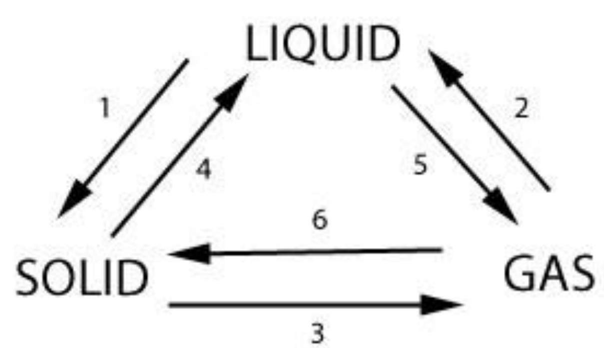
9. Which of the following would be the best choice as an insulator, gold, iron, aluminum or water? Explain
**Water; of the substances listed water has the highest specific heat capacity.**

10. Which of the following is the best conductor of heat, carbon, water, copper, or mercury? Explain.
**Mercury; of the substances listed, mercury has the lowest specific heat capacity.**

11. If a **q** value is positive, is a reaction endothermic or exothermic?
**If q is positive it is an endothermic change; if q is negative it is an exothermic change.**

**Part 2: Use the diagram to the bottom right to identify the process occurring. Indicate whether it is endothermic or exothermic.**

1. **<u>Freezing - Exothermic</u>**
2. **<u>Condensing - Exothermic</u>**
3. **<u>Sublimation - Endothermic</u>**
4. **<u>Melting - Endothermic</u>**
5. **<u>Vaporization - Endothermic</u>**
6. **<u>Deposition - Exothermic</u>**



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**Energy - Answers**

**Part 3: Solve each phase change problems in kilojoules and Calories (kilocalories). Indicate if an endothermic or exothermic change takes place.**

1. Calculate the amount of energy required to change 15.0 grams of ice at 0°C to 15.0 grams of water at 0°C. $H_{fus}$ = 80.0 cal/gram

**q = $H_{fus}$ x mass**
**q = 80.0 x 15.0**
**q = 1200 calories = 1.20 kilocalories = 5.02 kJ**
**endothermic**

2. If the heat of fusion of water is 340 J/g, calculate the amount of heat energy required to change 40.0 grams of ice at 0°C to 40.0 grams of water at 0°C.

**q = $H_{fus}$ x mass**
**q = 340 x 40.0**
**q = 13600 joules = 14 kilojoules = 3.3 kcal**
**endothermic**

3. Calculate the amount of energy released as 100.0 grams of steam are condensed into liquid water at 100°C.

**q = $H_{cond}$ x mass**
**q = -2259.0 x 100.0**
**q = -225900 joules or -225.9 kilojoules = -54.00 kcal**
**exothermic**

4. Calculate the amount of energy released as 35.0 grams of water at 0°C changes to 35.0 grams of ice at 0°C.

**q = $H_{sol}$ x mass**
**q = -335 x 35.0**
**q = -11725 joules or -11.7 kilojoules = -2.80 kcal**
**exothermic**

5. Calculate the amount of heat energy released as 220.0 grams of water at 0°C is changed to 220.0 grams of ice at 0°C.

**q = $H_{sol}$ x mass**
**q = -335 x 220.0**
**q = -73700 joules or -73.7 kilojoules = -17.6 kcal**
**exothermic**

6. Calculate the amount of energy absorbed as 400.0 grams of water is vaporized at 100°C.

**q = $H_{vap}$ x mass**
**q = 2259.0 x 400.0**
**q = 903600 joules or 903.6 kilojoules = 216.0 kcal**
**endothermic**

7. If the heat of fusion of FeO is 32.2 kJ/mol, calculate the amount of heat energy required to melt 17.7 grams of FeO.

**17.7 ÷ 71.8 = 0.247 moles**
**q = $H_{fus}$ x moles**
**q = 32.2 x 0.247**
**q = 7.95 kJ = 1.90 kcal**
**endothermic**

8. Calculate the kilojoules of energy needed to vaporize 70.0 grams of $NH_3$. $H_{vap}$ = 23.4 kJ/mol

**70.0 ÷ 17.03 = 4.12 moles**
**q = $H_{vap}$ x moles**
**q = 23.4 x 4.12**
**q = 96.4 kJ = 23.0 kcal**
**endothermic**

9. Calculate the amount of energy released as 55.0 grams of $Br_2$ is changed from liquid to solid. $H_{fus}$ = 10.72 kJ/mol

**55.0 ÷ 159.8 g = 0.344 moles**
**q = $H_{solid}$ x moles**
**q = -10.72 x 0.344**
**q = -3.69 kJ**
**exothermic**

10. The energy required to melt a solid into a liquid is called — **heat of fusion**

11. The energy released when freezing a liquid to a solid is called — **heat of solidification**

12. The energy released when condensing a gas to a liquid is called — **heat of condensation**

13. The energy required to vaporize a liquid into a gas is called — **heat of vaporization**

14. Heat of condensation is an (endothermic / **exothermic**) process. Circle one.
15. Heat of fusion is an (**endothermic** / exothermic) process. Circle one.
16. Heat of vaporization is an (**endothermic** / exothermic) process. Circle one.
17. Heat of solidification is an (endothermic / **exothermic**) process. Circle one.
18. Heat of fusion is used at the: (**melting point** / boiling point / freezing point / condensing point) Circle one.
19. Heat of vaporization is used at the: (melting point / **boiling point** / freezing point / condensing point) Circle one.
20. Heat of solidification is used at the: (melting point / boiling point / **freezing point** / condensing point) Circle one.
21. Heat of condensation is used at the: (melting point / boiling point / freezing point / **condensing point**) Circle one.

**Part 4: Multiple Choice Problems -** Choose the best answer for each of the following.

1. How much heat is required to raise the temperature of 100. g of $Fe_2O_3$ from 5.0°C to 25.0°C? (Specific heat $Fe_2O_3$, 0.634 $J \cdot g^{-1} \cdot °C^{-1}$)

(A) 1.58 kJ **(B) 1.27 kJ** (C) 0.845 kJ (D) 0.0634 kJ (E) 1.902 kJ

2. The specific heats of several metals are given in the table. If the same number of Joules were applied to the same mass of each metal, which metal would show the greatest temperature change?

| Substance | Specific Heat, $J \cdot g^{-1} \cdot °C^{-1}$ |
|---|---|
| Al | 0.900 |
| Au | 0.129 |
| Cu | 0.385 |
| Hg | 0.139 |

(A) Al **(B) Au** (C) Cu (D) Hg (E) all would be equal