

Practice Test for Thermochemistry ANSWERS

1. Calculate the number of joules of energy are needed to melt a 10.0 pound block of ice at zero degrees centigrade into water at the same temperature.

Convert pounds to grams first = 4540 grams of ice.

$$q = mH_f$$

$$q = (4540 \text{ g})(334 \text{ J/g}) = 1,516,360 \text{ J}$$

$$= 1,520,000 \text{ J (with 3 SF)}$$

2. Convert the number of joules you just calculated in question #1 into kilo-joules, calories, and Calories.

$$A \quad \frac{1,520,000 \text{ J}}{1} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 1520 \text{ kJ} \quad \text{with 3 SF}$$

$$B \quad \frac{1,520,000 \text{ J}}{1} \times \frac{1 \text{ cal}}{4.18 \text{ J}} = 363636.3636 \text{ cal} = 364,000 \text{ cal} \quad \text{with 3 SF}$$

$$C \quad \frac{364,000 \text{ cal}}{1} \times \frac{1 \text{ Calorie}}{1000 \text{ cal}} = 364 \text{ Calories (capital "C" food Calories)} \\ \text{with 3 SF}$$

3. Calculate how many joules of energy are needed to vaporize 4.54 kg of liquid water at 100°C to steam at the same temperature.

Convert kg to grams = 4540 grams

$$q = mH_v$$

$$q = (4540 \text{ g})(2260 \text{ J/g}) = 10,260,400 \text{ J}$$

$$q = 10,300,000 \text{ J} \quad \text{with 3 SF}$$

4. Convert the number of joules you just calculated in question #3 into kilo-joules, calories, and Calories.

$$A \quad \frac{10,300,000 \text{ J}}{1} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 10,300 \text{ kJ} \quad \text{with 3 SF}$$

$$B \quad \frac{10,300,000 \text{ J}}{1} \times \frac{1 \text{ cal}}{4.18 \text{ J}} = 246114.833 \text{ cal} = 264,000 \text{ cal} \quad \text{with 3 SF}$$

$$C \quad \frac{264,000 \text{ cal}}{1} \times \frac{1 \text{ Calorie}}{1000 \text{ cal}} = 264 \text{ Calories (capital "C" food Calories)} \\ \text{with 3 SF}$$

5. Calculate how much energy in joules is required to warm 73.2 g pure water at 17.0°C to 29.5°C.

$$\begin{aligned}q &= m\Delta T \\q &= (73.2 \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(29.5-17.0^\circ\text{C}) \\q &= (73.2 \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(12.5^\circ\text{C}) \\q &= 3827.4 \text{ J} \\q &= 3820 \text{ J} \quad \text{with 3 SF}\end{aligned}$$

6. Convert the number of joules you just calculated in question #5 into kilo-joules, calories, and Calories.

$$\text{A} \quad \frac{3820 \text{ J}}{1} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 3.82 \text{ kJ} \quad \text{with 3 SF}$$

$$\text{B} \quad \frac{3820 \text{ J}}{1} \times \frac{1 \text{ cal}}{4.18 \text{ J}} = 913.8755981 \text{ cal} = 914 \text{ cal} \quad \text{with 3 SF}$$

$$\text{C} \quad \frac{914 \text{ cal}}{1} \times \frac{1 \text{ Calorie}}{1000 \text{ cal}} = 0.914 \text{ Calories (capital "C" food Calories)} \quad \text{3 SF}$$

Which of each of these pairs of reactions is MOST ENDOTHERMIC?

A. One mole C_2H_2 synthesizes $\Delta H = +227.4 \text{ kJ}$ or two moles nitrogen monoxide synthesizes $\Delta H = +182.6 \text{ kJ}$

B. One mole NaCl ionizes in water $\Delta H = +3.88 \text{ kJ}$ or one mole ammonium chloride does the same $\Delta H = +14.78 \text{ kJ}$

C. Two moles NO_2 synthesizes $\Delta H = +66.4 \text{ kJ}$ or two moles KNO_3 ionizes in water $\Delta H = +34.89 \times 2 = 69.78 \text{ kJ}$

This was not a "trick" question, but a Stoichiometry matters question. Make sure you watch the mole ratios on table I, and use them, don't just be lulled into complacency.

8. If you dissolve 120. grams NaOH solid into 1.0 liters of water which is at 11.0°C, how much energy is released (in joules)?

First determine how many moles of NaOH you have, so you can use Stoichiometry with the Table I reaction.

$$\frac{120. \text{ g NaOH}}{1} \times \frac{1 \text{ mole NaOH}}{40 \text{ g NaOH}} = 3.00 \text{ moles NaOH}$$

Now let's take this number of moles sodium hydroxide to the next page, and use it.

Table I shows that when 1 mole NaOH dissolves into water it releases 44.51 kJ of energy, ($\Delta H = -44.51$ kJ)

$$\text{MR} \quad \frac{\text{NaOH}}{\text{energy}} \quad \frac{1}{44.51 \text{ kJ}} = \frac{3}{X \text{ kJ}} \quad \text{Solve for } X \quad X = 133.53 \text{ kJ} \quad X = 134 \text{ kJ}$$

Which means that when 3.00 moles NaOH dissolves into H_2O , it releases 134 kJ **with 3SF**.

9. Using the info in question number 8, if you did dissolve that NaOH into the water, what would the final temperature of the water be?

Since 134 kJ of energy gets released into the one liter of water, it's exothermic, so the water gets hotter. How much hotter is knowable, if you use the basic heat formula. One liter water = 1000 mL = 1000 grams of water. Water start temp is 11.0°C . Convert the kJ to Joules to put in as your q value. The ΔT is our unknown in this problem.

$$q = mC\Delta T \\ 134,000 \text{ J} = (1000 \text{ grams})(4.18 \text{ J/g}\cdot^\circ\text{C})(\Delta T)$$

$$\frac{134,000 \text{ J}}{(1000 \text{ grams})(4.18 \text{ J/g}\cdot^\circ\text{C})} = \Delta T$$

$$32.05741... = \Delta T$$

$$\Delta T = 32.1^\circ\text{C} \quad \text{with 3 SF}$$

And now we think more... Is this our answer? The question was what is the final temp of the water. We solved for the change in temp. The final temp would be done this way: start temp + change in temp (exo) yields final temp.
So, $11.0^\circ\text{C} + 32.1^\circ\text{C} = 43.1^\circ\text{C}$

10. How much energy is required to cool 712 grams of steam at 101.0°C to ice at -1.25°C ? (Note: $C_{\text{steam}} = 1.7 \text{ J/g}\cdot^\circ\text{C}$, while the $C_{\text{ice}} = 2.10 \text{ J/g}\cdot^\circ\text{C}$) This is a **5** step Thermochemistry problem, which exceeds the regents demands, but you are from Vestal, so you can do this. Add the total joules together for one final answer.

First we think this through. We need to

- A. cool the steam down to 100°C
- B. phase change the steam to liquid water with no temperature change
- C. cool the water down from 100°C to 0°C
- D. phase change the liquid water to ice with no temperature change
- E. cool the ice down from zero to -1.25°C

This takes five steps, all releasing joules of energy. Each step is easy, then we add up the five parts to get one whole answer. Go slowly, it's fun...

step	formula needed	math setup	part answer (all 3 SF)
A	$q = mC\Delta T$	$q = (712 \text{ g})(1.7 \text{ J/g}\cdot^\circ\text{C})(1.0^\circ\text{C})$	= 1210 J
B	$q = mH_v$	$q = (712 \text{ g})(2260 \text{ J/g})$	= 1,610,000 J
C	$q = mC\Delta T$	$q = (712 \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(100.0^\circ\text{C})$	= 298,000 J
D	$q = mH_f$	$q = (712 \text{ g})(334 \text{ J/g})$	= 238,000 J
E	$q = mC\Delta T$	$q = (712 \text{ g})(2.10 \text{ J/g}\cdot^\circ\text{C})(1.25^\circ\text{C})$	= 1870 J

Add up the five parts to arrive at the total number of Joules needed to let this steam starting at 101.0°C cool down to -1.25°C ice. Answer = 2,150,000 J (with 3SF)

11. You synthesize 1250 grams of aluminum oxide. How much energy is released, or absorbed? First determine how many moles of Al_2O_3 you have.

$$\frac{1250 \text{ g Al}_2\text{O}_3}{1} \times \frac{1 \text{ mole Al}_2\text{O}_3}{102 \text{ g Al}_2\text{O}_3} = 12.3 \text{ moles Al}_2\text{O}_3 \text{ with 3 SF}$$

Table I shows us that when 2 moles Al_2O_3 synthesize that $\Delta H = 3351 \text{ kJ}$, which means that energy is given off exothermically. Let's do some Stoichiometry.

$$\text{MR} \quad \frac{\text{Al}_2\text{O}_3}{\text{energy}} \quad \frac{2}{3351 \text{ kJ}} = \frac{12.3}{X \text{ kJ}} \quad \text{Solve for X} \quad \begin{array}{l} 2X = 41217.3 \text{ kJ} \\ X = 20608.65 \text{ kJ} \end{array}$$

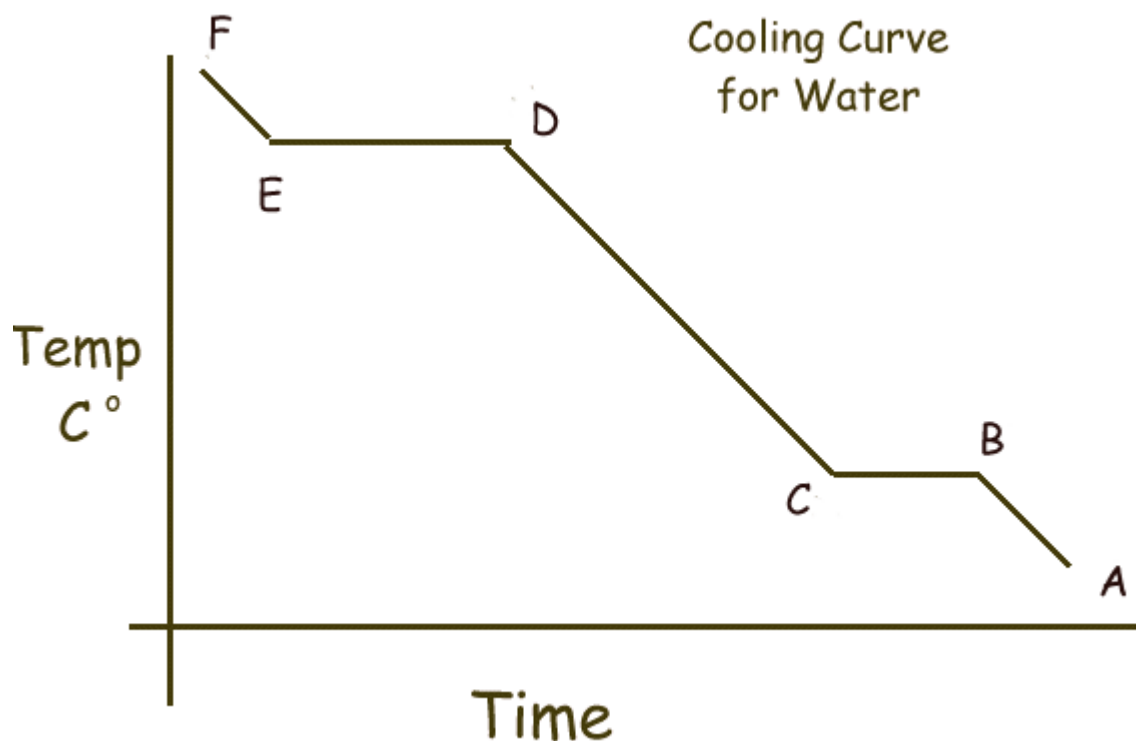
$X = 20,600 \text{ kJ with 3 SF}$

12. At which temperature does copper have the most kinetic energy?

A. 225 K B. 272 K C. 0°C D. 5°C

this is the highest temperature, which means the highest kinetic energy too.

Wow, that was a lot of work. Typing it all was really tough for me. It should get easier now...

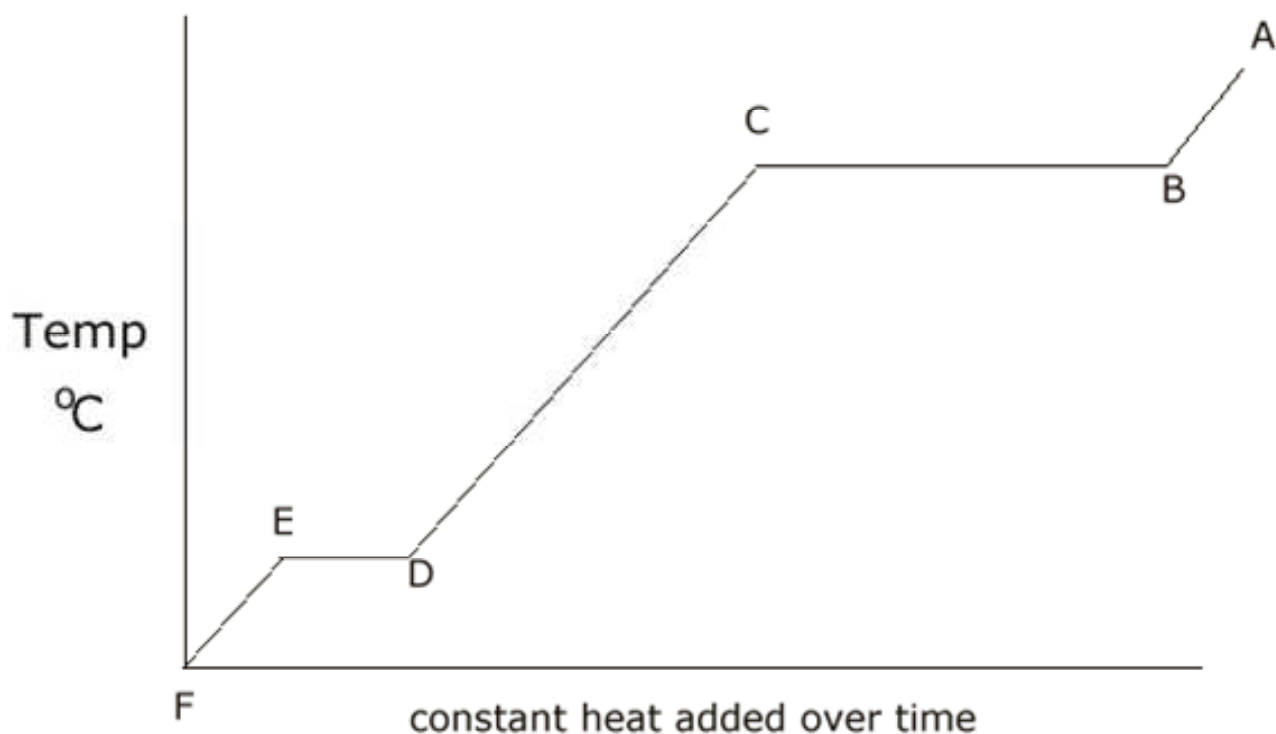


14. What is the temperature at CB? Does this temperature have a name?
 CB is the freezing/melting point, it's 0°C for water at standard pressure.
15. What is the temperature at ED? Does this temperature have a name?
 ED is the boiling/condensing point, it's 100°C for water at standard pressure.

16. If you were doing Thermochemistry problems, what formulas would you use for each segment listed in this chart? What unit for H₂O would you use?

segment	formula	water unit
FE	$q = mC\Delta T$	4.18 J/g·°C
ED	$q = mH_v$	2260 J/g
DC	$q = mC\Delta T$	4.18 J/g·°C
CB	$q = mH_f$	3334 J/g
BA	$q = mC\Delta T$	2.10 J/g·°C

Heating Curve for Water



17. What is the temperature at CB? Does this temperature have a name?
 CB is the boiling/condensing point, it's 100°C for water at standard pressure.
18. What is the temperature at ED? Does this temperature have a name?
 ED is the freezing/melting point, it's 0°C for water at standard pressure.

19. If you were doing Thermochemistry problems, what formulas would you use for each segment listed in this chart? What unit for H₂O would you use?

segment	formula	water unit
FE	$q = mC\Delta T$	2.10 J/g·°C
ED	$q = mH_F$	334 J/g
DC	$q = mC\Delta T$	4.18 J/g·°C
CB	$q = mH_V$	2260 J/g
BA	$q = mC\Delta T$	1.7 J/g·°C

20. How much energy is released from 35.0 grams of copper ($C = 0.39 \text{ J/g}\cdot^\circ\text{C}$) if it cools from 93.5 to 18.2 degrees centigrade?

Using the basic heat formula this is easy

$$q = mC\Delta T$$

$$q = (35.0 \text{ g})(0.39 \text{ J/g}\cdot^\circ\text{C})(75.3^\circ\text{C})$$

$$q = 1027.845 = 1030 \text{ J with 3 SF}$$

21. 52.3 g Hg changes temperature from 12.00°C to 33.25°C when 155.6 joules of heat are added. Calculate the specific heat capacity for mercury.

Using the basic heat formula this is easy

$$q = mC\Delta T$$

$$155.6 \text{ J} = (52.3 \text{ g})(C)(21.25^\circ\text{C})$$

solving for C

$$\frac{155.6 \text{ J}}{(52.3 \text{ g})(21.25^\circ\text{C})} = C$$

$$C = 0.1400 \text{ J/g}\cdot^\circ\text{C}$$

22. On both the heating curve & cooling curve, why is the line segment at 100°C so much longer than at 0°C? Use the numbers 334 and 2260 (with units) in your answer. How much difference is there between the two segments?

To phase change water on the hot side (100°C) you must add or remove 2260 joules per gram of water. On the cold phase change (at 0°C) it only takes 334 joules per gram to change phases. 2260 is about 7X bigger than 334, and therefore it takes about 7X more energy to vaporize than to melt water. It also says that steam condensing releases 7X more energy than when water freezes.

23. If $\text{O}_{2(\text{G})} \longrightarrow \text{O}_{(\text{G})} + \text{O}_{(\text{G})}$

Which best describes the process of this chemical reaction?

The bond breaks, and energy must be absorbed to do that, so...

A. energy is absorbed when a bond is broken

24. If $\text{I} + \text{I} \longrightarrow \text{I}_2$

Which statement best describes this chemical process?

The bond forms, which releases energy, so

B. a bond forms and energy is released

25. In a reaction on Table I you find that $\Delta H = +182.6 \text{ kJ}$. This reaction is

A. endothermic, the reaction absorbs energy

26. In a different reaction you find that $\Delta H = -571.6 \text{ kJ}$. This reaction is

D. exothermic, the reaction releases energy