

AP Chemistry

Thermochemistry

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Name: Key

Date: _____ Period: _____

1. Which "type" of system best describes a hot bowl of soup?

Open system: Both mass and energy can be exchanged with the surroundings

2. What is the change in internal energy of a system that has 682J of work done on it and gives off 485J of heat?

$$\left. \begin{array}{l} 682\text{J in (+)} \\ 485\text{J out (-)} \end{array} \right\} \text{net } \boxed{+197\text{J}}$$

3. When heat is absorbed by the system from the surroundings, the process is said to be endothermic, and the sign of q_{process} is positive.

4. Which is not true for an endothermic reaction?

- The temperature of the surroundings decreases.
- The enthalpy change for the reaction is positive.
- Heat flows from the surroundings into the system.
- The products have a lower enthalpy than the reactants. *enthalpy increases as energy is added to the system*
- All of the above are true.

5. What are the signs for q and w if a system absorbs 180J of heat energy while expanding against a constant pressure?

q is positive
w is negative $w = -P\Delta V$
 $\Delta V = V_f - V_i$
 $= \text{positive}$

6. At what velocity (m/s) must a 20.0g object be moving in order to possess a kinetic energy of 1.00J?

$$KE = \frac{1}{2}mv^2$$

$$1.00 = \frac{1}{2}(0.0200\text{kg})v^2 \quad 100 = v^2$$

$$\boxed{v = 10\text{m/s}}$$

7. What is the change in internal energy for a system that performs 213kJ of work on its surroundings and loses 79kJ of heat in the process?

$$\Delta E = q + w$$

$$q = -79\text{kJ}$$

$$w = -213\text{kJ}$$

$$= \boxed{-292\text{kJ}}$$

8. Which one of the following is false concerning internal energy, E ?

- It is a state function.
- It can be measured exactly. *Too many factors, but usually just interested in the change, ΔE*
- $\Delta E = E_{\text{final}} - E_{\text{initial}}$
- $\Delta E = q + w$
- When a system undergoes a process in which it gains energy from the surroundings, the ΔE for the process is positive.

9. Given: heat (q), internal energy (E), work (w) and enthalpy (H), which one(s) is/are state functions? What does it mean to be a state function?

E & H are state functions, while q & w are not. State functions only depend on the current state of the system, not its history.

10. The energy change in a system that occurs at a constant pressure is enthalpy $\Delta H = q_p$, while the energy change in a system that occurs at a constant volume is ΔE . (Use ΔH , ΔE and q_p as potential answers)

11. Explain the difference between ΔE and ΔH .

The difference is the $P\Delta V$ work

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

12. An object possesses a gravitational potential energy of 600J at a height of 10.0m. What will its kinetic energy, potential energy and velocity be at 5.00m when it is dropped? ($PE = mgh$, $KE = \frac{1}{2}mv^2$) *PE = mgh At 5.00m that is 50% of what it was*

$$PE = (50)(600\text{J}) = 300\text{J}$$

$$KE = 300\text{J}$$

$$300 = \frac{1}{2}(6)v^2$$

$$100 = v^2$$

$$\boxed{v = 10\text{m/s}}$$

$$m = 600 / (10.0 \times 10.0) = 6\text{kg}$$

13. Of the following types of energy, which are classified as kinetic energy?

1. thermal energy
2. mechanical energy
3. electrical energy

- a) 1 only b) 2 only
c) 3 only d) 2 and 3

e) 1, 2, and 3 See page 236

14. Which one of the following statements is INCORRECT?

- a) Energy is neither created nor destroyed in chemical reactions.
- b) Kinetic energy is the energy associated with motion.

- c) Exothermic processes transfer heat from the surrounding into the system.** *other way around*
d) Increasing the thermal energy of a gas increases the motion of its atoms.
e) Energy is the capacity to do work.

15. If the same amount of heat is added to 5.00 g samples of each of the metals below, which metal will experience the greatest temperature change?

Metal Specific Heat Capacity (J/g·K)

Al	0.897
Au	0.129
Cu	0.385
Fe	0.449
K	0.753

- a) Al **b) Au** *least energy needed to raise 1g by 1 K*
c) Cu d) Fe

16. Which of the following statements is CORRECT?

a) If a reaction occurs at constant pressure, $w = \Delta E$. *work = -pΔV*

b) If a reaction occurs at constant pressure, $q = \Delta H$.

c) If a reaction occurs at constant pressure, $q = \Delta E$. *q = ΔE + PΔV = ΔH*

d) If a reaction occurs at constant volume, $\Delta E > \Delta H$. $\Delta H = \Delta E + P\Delta V = \Delta E$ since $\Delta V = 0$

e) If a reaction occurs at constant volume, $w = \Delta E$. *At constant volume no work is done*

17. One statement of the first law of thermodynamics is that

a) the amount of work done on a system is dependent of pathway.

b) the total work done on a system must equal the heat absorbed by the system.

c) the heat flow in or out of a system is independent of pathway.

d) the total energy flow in or out of a system is equal to the sum of the heat transferred to or from the system and the work done by or on the system. *Consrv. of Energy*

e) in any chemical process the heat flow must equal the change in enthalpy.

18. A 2.885 g sample of methanol, CH_3OH , is combusted in a bomb calorimeter. The temperature of the calorimeter increases by 11.38 K. If the heat capacity of the bomb is 727.1 J/K and it contains 1.200 kg of water, what is the heat evolved per mole of methanol combusted? The specific heat capacity of water is 4.184 J/g·K and the molar mass of methanol is 32.04 g/mol.

- a) -65.41 kJ/mol
b) -91.89 kJ/mol
c) -634.5 kJ/mol

d) -726.5 kJ/mol *11.38K (727.1 J/K)*
e) -1.019×10^6 kJ/mol *1700. g (4.184 J/g·K) (11.38°C)*
2.885/32.04 = .0900 mol *57,136.704 J* *65,411 kJ*
65,411.02 J + 65,411 kJ = 130,822 J
130,822 J / .0900 mol = 1,453,577 J/mol = 1,453.577 kJ/mol

19. Which of the following chemical equations does not correspond to a standard molar enthalpy of formation? *= 726.8 kJ/mol*

a) $\text{Ca(s)} + \text{C(s)} + 3/2 \text{O}_2(\text{g}) \rightarrow \text{CaCO}_3(\text{s})$

b) $\text{C(s)} + 1/2 \text{O}_2(\text{g}) \rightarrow \text{CO(g)}$

c) $\text{H}_2(\text{g}) + 1/2 \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\lambda)$

d) $\text{N}_2(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g})$

e) $\text{SO}_2(\text{g}) + 1/2 \text{O}_2(\text{g}) \rightarrow \text{SO}_3(\text{g})$

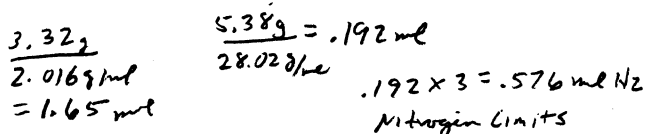
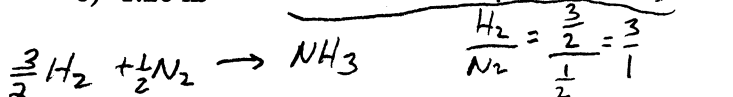
must form from elements

20. The standard molar enthalpy of formation of $\text{NH}_3(\text{g})$ is -45.9 kJ/mol. What is the enthalpy change if 5.38 g $\text{N}_2(\text{s})$ and 3.32 g $\text{H}_2(\text{g})$ react to produce $\text{NH}_3(\text{g})$?

a) -75.6 kJ b) -50.4 kJ

c) -17.6 kJ d) -8.81 kJ

e) -1.20 kJ



AP Chemistry

Thermochemistry 2

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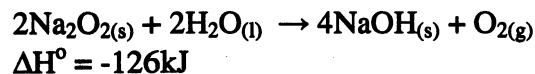
1. For the hypothetical reaction
 $3X_2(g) \rightarrow 2X_3(g)$, ΔH° is -333kJ (the atomic weight of X is 66.6amu). Determine ΔH of the conversion of 11.1g of X_3 gas into X_2 gas in kilojoules.

$$11.1\text{g } X_3 \left(\frac{1\text{ mol}}{199.8\text{g}} \right) = 0.0556\text{ mol } X_3 \left(\frac{+333\text{ kJ}}{2\text{ mol } X_3} \right)$$

9.25 kJ

Answer: 9.25 kJ

2. Given



Calculate the amount of heat that is evolved (in kJ) from the reaction of 25.0g of Na_2O_2 with water.

$$25.0\text{g } \text{Na}_2\text{O}_2 \left(\frac{1\text{ mol}}{77.98\text{g}} \right) = 0.321\text{ mol } \text{Na}_2\text{O}_2$$

$$0.321\text{ mol } \text{Na}_2\text{O}_2 \left(\frac{-126\text{ kJ}}{2\text{ mol } \text{Na}_2\text{O}_2} \right) = \underline{-20.2\text{ kJ}}$$

Answer: 20.2 kJ evolved

3. If the molar heat capacity of a compound with the formula $\text{C}_2\text{H}_6\text{SO}$ is $88.0\text{J/mol}\cdot\text{K}$, what is the specific heat (in $\text{J/g}\cdot\text{K}$) of this substance?

$$88.0\text{ J/mol}\cdot\text{K} \left(\frac{1\text{ mol}}{78.138\text{g}} \right) = \underline{1.13\text{ J/g}\cdot\text{K}}$$

Answer: 1.13 J/g·K

4. A sample of aluminum ($0.90\text{J/g}\cdot\text{K}$) absorbs 9.86J of energy while being heated from 23.2°C to 30.5°C . Determine the mass of the sample.

$$\Delta E = mC\Delta T$$

$$9.86\text{ J} = (m)(0.90)(30.5 - 23.2) = (0.90)(7.3)m$$

$$m = \underline{1.50\text{ g}}$$

Answer: 1.50 g

5. An iron sphere of diameter 4.00cm and density 7.86g/cm^3 at 20.0°C is dropped into a perfectly insulated beaker containing 90.0g of H_2O at 50.0°C ($V = 4/3\pi r^3$). The specific heats of Fe and H_2O are 0.45 and $4.18\text{J/g}\cdot\text{K}$, respectively. What is the final temperature (in $^\circ\text{C}$) of the iron sphere?

$$M_{\text{sphere}} = \left(\frac{4}{3}\pi \right) (2.00)^3 \cdot 7.86\text{ g/cm}^3$$

$$= 263.39\text{ g} \approx 263\text{ g}$$

$$M_{\text{H}_2\text{O}} C_{\text{H}_2\text{O}} \Delta T_{\text{H}_2\text{O}} + M_{\text{Fe}} C_{\text{Fe}} \Delta T_{\text{Fe}} = (90.0\text{ g})(4.18\text{ J/g}\cdot\text{K})(T_f - 50.0^\circ\text{C}) + (263.39\text{ g})(0.45\text{ J/g}\cdot\text{K})(T_f - 20.0^\circ\text{C}) = 0$$

Answer: 42.8 °C

6. When 50.0mL of 1.00M HCl is mixed with 50.0mL of 1.00M NaOH (both at 23.0°C), the resulting solution increases in temperature to 29.8°C . Assuming that the solution has the density and specific heat of pure water, calculate the enthalpy of the reaction.

$$1.00\text{ g/mL} \rightarrow 50.0\text{ mL}(1.00\text{ g/mL}) = 50.0\text{ g HCl}$$

$$\text{NaOH} \rightarrow 50.0\text{ mL}(1.00\text{ g/mL}) = 50.0\text{ g NaOH}$$

$$\Delta H_{\text{rxn}} = q_{\text{HCl}} + q_{\text{NaOH}} = (50.0)(4.184)(29.8 - 23) + (50.0)(4.184)(29.8 - 23)$$

$$= 2845\text{ J} = \underline{2.85\text{ kJ}}$$

Answer: 2.85 kJ

7. A 1.96g sample of titanium was burned in a bomb calorimeter that had a heat capacity of 9.84kJ/K . The temperature of the calorimeter increased from 36.84°C to 98.82°C . Calculate the heat (kJ) that is released from the combustion of one mole of titanium.

$$q_{\text{bomb}} = (9.84\text{ kJ/K})(98.82 - 36.84) = 609.9\text{ kJ}$$

$$609.9\text{ kJ} / 1.96\text{ g Ti} \left(\frac{47.867\text{g}}{1\text{ mol}} \right) = \underline{1.489 \times 10^4\text{ kJ}}$$

Answer: $1.489 \times 10^4\text{ kJ/mol Ti}$

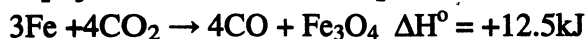
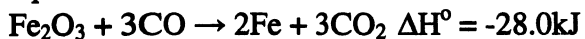
8. The heat of solution (ΔH_{soln}) for sodium hydroxide is 44.4kJ/mol . If a 13.9g sample of NaOH dissolves in $250.\text{g}$ of water at 23.0°C , what will the final temperature of the solution be? Assume that the solution has the same specific heat as water.

$$44.4\text{ kJ/mol} \left(\frac{1\text{ mol}}{39.998\text{g}} \right) (13.9\text{ g}) \left(\frac{1000\text{ J}}{1\text{ kJ}} \right) = 15430.\text{ J}$$

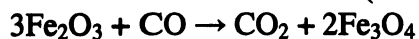
$$15430.\text{ J} = (263.9)(4.184)(x - 23.0) \Rightarrow x = \underline{36.97^\circ\text{C}}$$

Answer: 36.97 °C

9. Consider the following thermochemical equations:



Calculate the value of ΔH° (in kJ) for



$$\begin{aligned} &= 3(\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2) + \\ &2(3\text{Fe} + 4\text{CO}_2 \rightarrow 4\text{CO} + \text{Fe}_3\text{O}_4) \\ &= 3(-28.0\text{ kJ}) + 2(12.5\text{ kJ}) = \boxed{-59.0\text{ kJ}} \end{aligned}$$

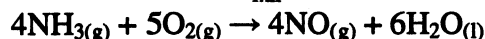
Answer: -59.0 kJ

10. For which of the species in the reaction below is the ΔH_f° equal to zero (more than one choice is possible)?



Answer: Co, H₂

11. Calculate $\Delta H_{\text{rxn}}^\circ$ for



given

Substance	ΔH_f° (kJ/mol)
H ₂ O(l)	-286
NO(g)	90
NH ₃ (g)	-46

$$\begin{aligned} &= [4(90) + 6(-286)] - [4(-46)] \\ &= \boxed{-1172\text{ kJ}/4\text{ mol NH}_3} \end{aligned}$$

Answer: -1172 kJ/4 mol NH₃

12. Given that silver has a specific heat of

0.235 J/g°C, what is this value in cal/g°C? If a

75.0g sample of silver at 100.°C is dumped into a

calorimeter containing 100.g of water at 20.0°C,

what will be the final temperature of the system

at equilibrium? How much heat energy is

transferred to the water in joules?

a) $0.235\text{ J/g}^\circ\text{C} \left(\frac{1\text{ cal}}{4.184\text{ J}}\right) = \boxed{0.0562\text{ cal/g}^\circ\text{C}}$

b) $q_{\text{H}_2\text{O}} + q_{\text{Ag}} = (m_{\text{H}_2\text{O}})(C_{\text{H}_2\text{O}})(\Delta T) + (m_{\text{Ag}})(C_{\text{Ag}})(\Delta T_{\text{Ag}})$
 $= (100.)(4.184)(T_f - 20.0^\circ\text{C}) + (75.0)(0.235)(T_f - 100^\circ\text{C})$
 $= 0 \Rightarrow \boxed{T_f = 23.2^\circ\text{C}}$

Answer: 0.0562 cal/g°C

Answer: 23.2 °C
 $q_{\text{H}_2\text{O}} = (100.)(4.184)(23.2 - 20.0) = \boxed{1353\text{ J}}$

Answer: 1350 J
 ↑ sig figs

13. A sample of a block of metal measuring

1.351cm x 10.00cm x 1.000cm (and density =

7.874g/cm³) at 100.°C is placed into a coffee cup

calorimeter containing 100.g of water at 22.0°C.

If the final temperature of the system is 30.0°C,

determine which of the following metals the

sample is likely to be.

Substance	Specific heat (J/g°C)
Aluminum	0.902
Iron	0.449
Lead	0.128
Silver	0.235

$$m_{\text{block}} = (1.351 \cdot 10.00 \cdot 1.000)\text{cm}^3 (7.874\text{ g/cm}^3) = 106.37\text{ g}$$

$$q_{\text{H}_2\text{O}} + q_{\text{block}} = (100\text{ g})(4.184\text{ J/g}^\circ\text{C})(30.0 - 22.0) + (106.37\text{ g})(x)(30.0 - 100)$$

 $= 0 \Rightarrow x = 0.4495 \Rightarrow \text{iron}$

Answer: iron

14. A 55.0g sample of water at 69.0°C is added

to an unknown amount of water at 22.0°C. The

final temperature is 31.0°C. How many grams of

water were there initially at 22.0°C?

$$0 = (55.0)(4.184)(31.0 - 69.0) + (m)(4.184)(31.0 - 22.0)$$

 $m = \boxed{232.9\text{ H}_2\text{O}}$

Answer: 232.9 g H₂O