

Name: _____

Date: _____

§ 16.01a

Thermochemistry: Calorimetry**Problems** (specific heat of H₂O = 4.184 J/(g °C))

1. What is specific heat? What are the units for it? Is it an intensive or extensive property?

- ① Specific heat = amount of heat required to raise the temperature of 1 gram of a substance by 1°C
- (A) temperature of 1 gram of a substance by 1°C
- (B) units = J/g.°C (J = Joules; 1 calorie = 4.184 J)
- (C) it is an intensive property & is used to identify an unknown substance.

2. Consider the following data:

Metal	Al	Cu
Mass (g)	10	30
Specific Heat (J/(g.°C))	0.900	0.385
Temperature (°C)	40	60

When of these two metals are placed in contact, which of the following will take place?

- (a) Heat will flow from Al to Cu because Al has a larger specific heat.
- (b) Heat will flow from Cu to Al because Cu has a larger specific heat.
- (c) Heat will flow from Cu to Al because Cu has a larger heat capacity.
- (d) Heat will flow from Cu to Al because Cu is at a higher temperature.

② $T_{Cu} > T_{Al}$ so heat flows from Cu to Al

The temperature change will be more for the copper because it has a lower specific heat and, thus, takes less energy to change its temp.

3. Consider two metals A and B, each having a mass of 100 g and an initial temperature of 70°C. The specific heat of A is larger than that of B. Under the same heating conditions, which metal would take longer to reach a temperature of 21°C?

③ "A" has a higher specific heat so it will take longer — it requires more energy to increase its temperature.

4. A 352-gram piece of silver at 100°C is placed in 100-grams of water at 25°C. The final temperature of the water is 37.60°C. What is the specific heat of silver?

	MASS (g)	TEMP (°C)		ΔT	S. (J/g·°C)
		Init	Final		
Ag	352	100	37.60	-62.40	?
H ₂ O	100	25	37.60	12.60	4.184

$$(m_{\text{Ag}})(s_{\text{Ag}})(\Delta T_{\text{Ag}}) = -(m_{\text{H}_2\text{O}})(s_{\text{H}_2\text{O}})(\Delta T_{\text{H}_2\text{O}})$$

$$(352)(s_{\text{Ag}})(-62.40) = -(100)(4.184)(12.60)$$

$$-21,965(s_{\text{Ag}}) = -5272$$

$$s_{\text{Ag}} = \frac{0.240 \text{ J/g}\cdot\text{°C}}$$

(agrees with textbook value)

5. A 28.4 g sample of aluminum is heated to 39.4°C, and then is placed in a calorimeter containing 50.0 g of water. The temperature of water increased from 21.00°C to 23.00°C. What is the specific heat of aluminum?

	MASS (g)	T (°C)		ΔT	S. (J/g·°C)
		Init	Final		
Al	28.4	39.4	23.00	-16.4	s_{Al}
H ₂ O	50.0	21.00	23.00	2.00	4.184 J/g·°C

$$(m_{\text{Al}})(s_{\text{Al}})(\Delta T_{\text{Al}}) = -(m_{\text{H}_2\text{O}})(s_{\text{H}_2\text{O}})(\Delta T_{\text{H}_2\text{O}})$$

$$s_{\text{Al}} = \frac{(m_{\text{H}_2\text{O}})(s_{\text{H}_2\text{O}})(\Delta T_{\text{H}_2\text{O}})}{(m_{\text{Al}})(\Delta T_{\text{Al}})}$$

$$s_{\text{Al}} = \frac{-(50.0)(4.184)(2.00)}{(28.4)(-16.4)}$$

$$s_{\text{Al}} = \underline{0.898 \text{ J/g}\cdot\text{°C}} \quad (\text{agrees with textbook value})$$

6. A sheet of gold weighing 10.0 g and at a temperature of 18.0°C is placed flat on a sheet of iron weighing 20.0 g and a temperature of 55.6°C. What is the final temperature of the combined metals? Assume no heat is lost to the surroundings. (*Hint*: The heat gained by the gold must be equal to the heat lost by the iron. The specific heats of gold and iron are 0.129 J/g·°C and 0.444 J/g·°C, respectively.)

	Mass (g)	T_{init} (°C)	T_{final} (°C)	ΔT	S
Au	10.0	18.0	T_f	ΔT_{Au}	0.129
Fe	20.0	55.6	T_f	ΔT_{Fe}	0.444

$$(m_{\text{Au}})(S_{\text{Au}})(\Delta T_{\text{Au}}) = (m_{\text{Fe}})(S_{\text{Fe}})(\Delta T_{\text{Fe}})$$

$$(10.0)(0.129)(T_f - 18.0) = -(20.0)(0.444)(T_f - 55.6)$$

$$1.29(T_f - 18.0) = -(8.88)(T_f - 55.6)$$

$$1.29T_f - 23.22 = -8.88T_f + 493.8$$

$$\Rightarrow +8.88T_f + 23.22 + 8.88T_f + 23.22$$

$$1.29T_f + 8.88T_f = 493.8 + 23.22$$

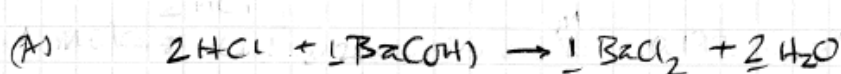
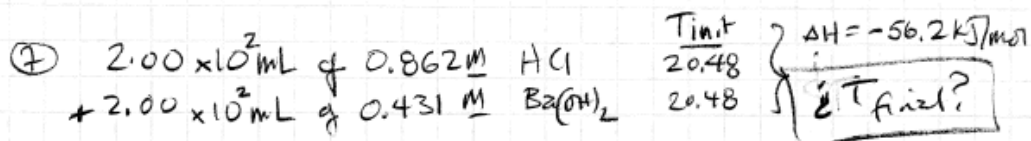
$$10.17T_f = 517.0$$

$$\Rightarrow \frac{10.17T_f}{10.17} = \frac{517.0}{10.17}$$

$$T_f = 50.84^\circ\text{C}$$

Extension

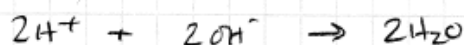
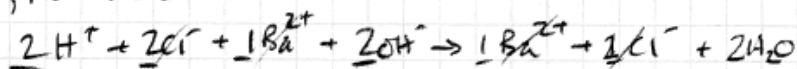
7. A quantity of 2.00×10^2 mL of 0.862 M HCl is mixed with 2.00×10^2 mL of 0.431 M $\text{Ba}(\text{OH})_2$ in a constant-pressure calorimeter of negligible heat capacity. The initial temperature of the HCl and $\text{Ba}(\text{OH})_2$ solutions is the same at 20.48°C . For the process: $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$, the heat of neutralization is -56.2 kJ/mol. What is the final temperature of the mixed solution?



(B) moles: HCl: $(0.200 \text{ L})(0.862 \frac{\text{mol}}{\text{L}}) = 0.172 \text{ mol HCl} = 0.172 \text{ mol H}^+$

$\text{Ba}(\text{OH})_2 (0.200 \text{ L})(0.431 \frac{\text{mol}}{\text{L}}) = 0.0862 \text{ mol Ba}(\text{OH})_2$
 $\rightarrow 0.172 \text{ mol OH}^-$

(C) rxn, net ionic:



So 0.172 mol H^+ neutralized 0.172 mol OH^-
 to form $0.172 \text{ mol H}_2\text{O}$

(D) Energy (J): $0.172 \text{ mol H}_2\text{O} \times \frac{-56.2 \text{ kJ}}{1 \text{ mol H}_2\text{O}} = 9.667 \text{ kJ released}$
 (-9.667 kJ)

(E) volume \Rightarrow mass of solution: $200 \text{ mL} + 200 \text{ mL} = 400 \text{ g}$

$q = (m)(s)(\Delta T)$

$9.667 \text{ kJ} = (400 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(T_{\text{final}} - 20.48^\circ\text{C})$

(9.667 J)

$9,667 = (1,673.6 \text{ J})(T_f - 20.48^\circ\text{C})$

$9,667 = 1,673.6 T_f + \frac{-34,275}{+34,275}$

$\frac{43,943}{1,673.6} = \frac{1,673.6 T_f}{1,673.6}$

$26.26^\circ\text{C} = T_f$

8. A 0.1375-g sample of solid magnesium is burned in a constant-volume bomb calorimeter that has a heat capacity of $3024 \text{ J/}^\circ\text{C}$. The temperature increases by 1.126°C . Calculate the heat given off by the burning Mg, in kJ/g and kJ/mol.