Name: Date:

§ 16.01a **Thermochemistry: Calorimetry**

Problems (specific heat of H_2O = 4.184 J/(g ^oC)

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
(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.^{o}C))	0.900	0.385
Temperature (^{o}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^{o}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^{o}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) Init	Final	ΔT	S (J/g·°C)
Ag	352	100	37.60	−62.40	?
H_2O	100	25	37.60	12.60	4.184

$$(m_{Ag})(S_{Ag})(\Delta T_{Ag}) = -(m_{H_2O})(S_{H_2O})(\Delta T_{H_2O})$$

$$(352)(S_{Ag})(-62.40) = -(100)(4.184)(12.60)$$

$$-21{,}965(S_{Ag}) = -5272$$

$$S_{Ag} = \underline{0.240\ J/g\cdot°C}$$

(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) Init	T_{final}	ΔT	S (J/g°C)
Al	28.4	39.4	23.00	−16.4	S_{Al}
H_2O	50.0	21.00	23.00	2.00	4.184 J/g°C

$$(m_{Al})(S_{Al})(\Delta T_{Al}) = -(m_{H_2O})(S_{H_2O})(\Delta T_{H_2O})$$

$$S_{Al} = \frac{-(m_{H_2O})(S_{H_2O})(\Delta T_{H_2O})}{(m_{Al})(\Delta T_{Al})}$$

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

$$S_{Al} = \underline{0.898\ J/°C\cdot g}$$ (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.

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

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

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

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

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

$$\Rightarrow +8.88\,T_f + 23.22 \quad +8.88\,T_f \quad + 23.22$$

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

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

$$T_f = 50.84^{o}C$$

Extension

7. A quantity of 2.00×10^2 mL of 0.862 HCl *M* is mixed with 2.00×10^2 mL of 0.431 *M* $Ba(OH)_2$ in a constant-pressure calorimeter of negligible heat capacity. The initial temperature of the HCl and $Ba(OH)_2$ solutions is the same at 20.48°C. For the process: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$, the heat of neutralization is –56.2 kJ/mol. What is the final temperature of the mixed solution?

(7) 2.00×10^2 mL of 0.862 M HCl   Tinit 20.48 $\Delta H = -56.2$ kJ/mol
+ 2.00×10^2 mL of 0.431 M $Ba(OH)_2$ 20.48 $\boxed{¿T_{final}?}$

(A) $2HCl + 1Ba(OH) \rightarrow 1BaCl_2 + 2H_2O$

(B) moles: HCl: (0.200 L)(0.862 mol/L) = 0.172 mol HCl = 0.172 mol H^+

$Ba(OH)_2$ (0.200 L)(0.431 mol/L) = 0.0862 mol $Ba(OH)_2$
→ 0.172 mol OH^-

(C) rxn, net ionic:

$2H^+ + 2Cl^- + 1Ba^{2+} + 2OH^- \rightarrow 1Ba^{2+} + 1Cl^- + 2H_2O$

$2H^+ + 2OH^- \rightarrow 2H_2O$

so 0.172 mol H^+ neutralized 0.172 mol OH^-
to form 0.172 mol H_2O

(D) Energy (J): $\frac{0.172 \text{ mol } H_2O}{1} \times \frac{-56.2 \text{ kJ}}{1 \text{ mol } H_2O}$ = 9.667 kJ released
(–9.66 kJ)

(E) volume → mass of solution: 200 mL + 200 mL = 400 g

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

9.667 kJ = (400 g)(4.184 J/g°C)(T_{final} – 20.48 °C)
↓
(9,667 J)

9,667 = (1,673.6 J)(T_f – 20.48°C)
9,667 = 1,673.6 T_f + –34,275
+34,275   +34275

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

$\boxed{26.26°C = T_f}$

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