Phase Change

① bag i: 220.0 g ice, T_i 0°C — bag f: 220.0°C, $T_f = 21.0°C$

$q_{ice\ phase\ \Delta} + q_{ice\ warm} = $ total heat

$+ \frac{333 J}{g} \times 220 g + \frac{4.184 J}{g°C} \times 220.0 g \times (21.0°C - 0.0°C)$

↑ melting endothermic

73,260 J + 19,330 J = 92,590 J

$q_{total} = 92.6$ kJ added

② $EtOH_{(s)}$ −180.3°C →① $EtOH_{(s)}$ −117.3°C →② $EtOH_{(l)}$ −117.3°C →③ $EtOH_l$ 78.3°C →④ $EtOH_{(g)}$ 78.3°C

$\frac{0.971 J}{g°C} \times 10.0 g \times (-117.3 - (-180.3))°C + \frac{218 J}{g} \times 10.0 g +$ ~~$\frac{230 J}{g°C}$ (10.0 g)~~

(+)

$\frac{2.30 J}{g°C} \times 10.0 g \times (78.3°C - (-117.3°C)) + \frac{854 J}{g} \times 10.0 g$

611.73 J + 2180 J + 4498.8 J + 8540 J

$q_T = + 15830.53 J$

15830 J → +15.83 kJ

Reactions in Calorimeters

(1)

50.0 mL 0.100 M $AgNO_3$ 50.0 mL 0.100 M NaCl

T_i = 22.60°C 100.0 mL total volume

T_f = 23.40°C = 100.0 g total mass

$q_{Rxn} = -q_{sol}$

$= 4.184 \frac{J}{g°C} \times 100.0 g \times (23.40°C - 22.60°C)$

= 334.4 J ∴ q_{Rxn} = −334.4 J

$\Delta H_{Rxn} = \frac{-334.4 J}{0.01 \text{ mol AgCl}}$ = 33440 J/mol = 33 kJ/mol AgCl

The answer should be –334.72 J, this is divided by 0.0500mol AgCl. ΔH = –66944 J/mol. To the correct SF, ΔH = -67kJ/mol

Reactions in Calorimeters

Limiting Reagent: $2HCl_{aq} + Ba(OH)_2 \rightarrow BaCl_{2aq} + 2H_2O_l$

	HCl	$Ba(OH)_2$
	100.0 mL	300.0 mL
	0.500 M	0.500 M
	50.0 mmol HCl	150. mol $Ba(OH)_2$
	LR	

a) $\Delta H_{Rxn} = \frac{-118\ kJ}{2\ mol\ HCl} \times 50.0\times10^{-3}\ mol\ HCl = -2.95\ kJ$ released

b) The initial and final masses are the same.

~~$q_{initial} = q_{final}$~~

$$-4.18\frac{J}{g^\circ C} \times 400.g \times (T_f - 25.00^\circ C) = -2.95\ kJ$$

$$-1673.6T_f + 41{,}840J = -2.95\ kJ$$

$$-1673.6\frac{J}{^\circ C}\ T_f = -44{,}790J$$

$$T_f = 26.72^\circ C \Rightarrow 26.7^\circ C$$

Warming and Cooling

① $q_{metal} + q_{H_2O} = 0$

$$Cp_{metal} \cdot mass_{metal} \cdot \Delta T_{metal} = -Cp_{H_2O} \cdot mass_{H_2O} \cdot \Delta T_{H_2O}$$

$$Cp_{metal} = \frac{-Cp_{H_2O} \cdot mass_{H_2O} \cdot \Delta T_{H_2O}}{mass_{metal} \cdot \Delta T_{metal}}$$

$$\frac{-\left(4.184 \frac{J}{g°C} \times 75.0g\ H_2O \times (28.34°C - 24.00°C)\right)}{26.00g \times (28.34°C - 82.25°C)}$$

$$= -1361.9J / -1401.66g°C = 0.972 \frac{J}{g°C}$$

The problem was misleading. Water in the liquid phase does not exisst at 1 atm, and 135°C. It is now steam. q(steam) provides the energy to warm the 95 degree water, and possibly allow a phase change, but not all of the water can go to steam.

q steam = -q water
50.0 g x 1.841J/g°Cx–35°C = –3224J. this is the amount of energy the 95°C water absorbs. 3222J = 4.184J/g°Cx70.0 g x 5°C + energy for phase change
3222J–1464J = 1758 J remaining. ONlY much energy is absorbed for the phase change which equates to 0.779g water.
it is the only mass of the 95°C water that can possibly have a T higher than 100°C if there is enough energy to change the temperature.
alas, there is not.
50.0 g x 1.841J/g°C x (Tf-100) = –0.779g x 1.841J/g°C(Tf-100), Tf = 100. there is not enough energy to raise the temperature dramatically past 100°C for the steam. all 50.779 g of steam is at 100°C

Bomb Calorimetry

(1) $-q_{cal} = q_{rxn}$

$\Delta H_{comb} \cdot \text{mass cmp'd} = -C_v \cdot \Delta T_{cal}$

$\frac{-5903\ kJ}{mol} \Big| \frac{1\ mol}{152.23\ g} \Big| 0.1204\ g\ \text{camphor} = -4.669\ kJ$

$C_v = \frac{-4.699\ kJ}{-2.23\ °C} = 2.05\ kJ/°C$

(2) $\Delta H_{comb} \cdot 0.1964\ g = -1.56\ \frac{kJ}{°C} \times 3.2\ °C$

$\Delta H_{comb}\ (\text{grams}) = \frac{-1.56\ \frac{kJ}{°C} \times 3.2\ °C}{0.1964\ g}$

$\approx -25\ kJ/g$

$\Delta H_{comb}\ (mol) = \frac{-25.4\ kJ}{g} \Big| \frac{108.1\ g}{mol} = -2{,}700\ kJ/mol$

Bomb Calorimeter

③ $C_v = ?$

$\Delta H_{comb} \cdot mass = -C_v \Delta T$

$$\frac{-26.42 \frac{kJ}{g} \times 0.1584 g}{-2.54 °C} = 1.647 \frac{kJ}{°C}$$

ΔH_{comb} vanillin

$$\Delta H_{comb} = \frac{-\left(1.647 \frac{kJ}{°C}\right)(3.25°C)}{0.2130 \text{ g vanillin}} = -25.139 \text{ kJ/g}$$

-25.1 kJ/g