Problems of the Day up to 10/30/09

The following answers are cut and pasted from the publisher's answer keys. Any additional commentary by me is in {} brackets.

- **2.90** (a) This is an ionic compound. Prefixes are *not* used. The correct name is barium chloride.
 - (b) Iron has a +3 charge in this compound. The correct name is iron(III) oxide.
 - (c) NO_2^- is the nitrite ion. The correct name is cesium nitrite.
 - (d) Magnesium is an alkaline earth metal, which always has a +2 charge in ionic compounds. The roman numeral is not necessary. The correct name is magnesium bicarbonate.
- 2.92 (a) Ammonium is NH_4^+ , not NH_3^+ . The formula should be $(NH_4)_2CO_3$.
 - (b) Calcium has a +2 charge and hydroxide has a -1 charge. The formula should be Ca(OH)₂.
 - (c) Sulfide is S^{2-} , not SO₃²⁻. The correct formula is CdS.
 - (d) Dichromate is $Cr_2O_7^{2-}$, not $Cr_2O_4^{2-}$. The correct formula is $ZnCr_2O_7$.

3.70 ? mol C = 7.86 g CO₂ × $\frac{1 \text{molCO}_2}{44.01 \text{g CO}_2}$ × $\frac{1 \text{mol C}}{1 \text{mol CO}_2}$ = 0.1786 mol C × $\frac{12.01 \text{g C}}{1 \text{mol C}}$ = 2.145 g C ? mol H = 2.14 g H₂O × $\frac{1 \text{molH}_2\text{O}}{18.02 \text{g H}_2\text{O}}$ × $\frac{2 \text{mol H}}{1 \text{mol H}_2\text{O}}$ = 0.2375 mol H $\frac{1.008 \text{g H}}{1 \text{mol H}}$ = 0.2394 g H Mass O = mass Sample - (Mass C + Mass H) = 2.856 g O × $\frac{1 \text{molO}}{16.00 \text{g O}}$ = 0.1785 mol O Molar ratios: C : H : O = 0.1786 : 0.2375 : 0.1785 = $\frac{0.1786}{0.1785}$: $\frac{0.2375}{0.1785}$ = 1.00 : 1.33 : 1.00

Multiplying all three numbers by 3 give the empirical formula $C_3H_4O_3$, which has an empirical formula weight of 88.06 g. {this is half the given molar mass of 176 g/mol} The molecular formula is $C_6H_8O_6$.

{When the hydrocarbon being burned also contains oxygen, the amount of O can only be determined indirectly. Once you find out the moles of C and H present in the sample, you can determine their masses and subtract them from the mass of the initial sample. The difference in mass is the oxygen.}

3.102 This is a limiting reactant problem. Calculate the moles of Li₃N produced assuming complete reaction for each reactant.

$$6\mathrm{Li}(s) + \mathrm{N}_2(g) \rightarrow 2\mathrm{Li}_3\mathrm{N}(s)$$

12.3 g Li ×
$$\frac{1 \text{ mol Li}}{6.941 \text{ g Li}}$$
 × $\frac{2 \text{ mol Li}_{3N}}{6 \text{ mol Li}}$ = 0.5907 mol Li₃N
33.6 g N₂ × $\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2}$ × $\frac{2 \text{ mol Li}_3N}{1 \text{ mol N}_2}$ = 2.398 mol Li₃N

Li is the limiting reactant; it limits the amount of product produced. The amount of product produced is 0.5907 mole Li₃N. Convert this to grams.

? g Li₃N = 0.5907 mol Li₃N ×
$$\frac{34.833 \text{ g Li}_3\text{N}}{1 \text{ mol Li}_3\text{N}}$$
 = **20.6 g Li₃N**

This is the theoretical yield of Li₃N. The actual yield is given in the problem (5.89 g). The percent yield

is:

% yield =
$$\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{5.89 \text{ g}}{20.6 \text{ g}} \times 100\% = 28.6\%$$

5.38 The heat gained by the calorimeter is:

$$q = C_{\rm p}\Delta T$$

 $q = (3024 \text{ J/°C})(1.126^{\circ}\text{C}) = 3.405 \times 10^3 \text{ J}$

The amount of heat given off by burning Mg in kJ/g is:

$$(3.405 \times 10^3 \text{ J}) \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{1}{0.1375 \text{ g Mg}} = 24.76 \text{ kJ/g Mg}$$

The amount of heat given off by burning Mg in kJ/mol is:

$$\frac{24.76 \text{ kJ}}{1 \text{ g Mg}} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 601.9 \text{ kJ/mol Mg}$$

5.39 Strategy: The neutralization reaction is exothermic. 56.2 kJ of heat are released when 1 mole of H^+ reacts with 1 mole of OH^- . Assuming no heat is lost to the surroundings, we can equate the heat lost by the reaction to the heat gained by the combined solution. How do we calculate the heat released during the reaction? Are we combining 1 mole of H^+ with 1 mole of OH^- ? How do we calculate the heat absorbed by the combined solution?

Solution: Assuming no heat is lost to the surroundings, we can write:

 $q_{\rm soln} + q_{\rm rxn} = 0$ $q_{\rm soln} = -q_{\rm rxn}$

or

First, let's set up how we would calculate the heat gained by the solution,

 $q_{\rm soln} = m_{\rm soln} s_{\rm soln} \Delta T$

where *m* and *s* are the mass and specific heat of the solution and $\Delta T = T_f - T_i$.

We assume that the specific heat of the solution is the same as the specific heat of water, and we assume that the density of the solution is the same as the density of water (1.00 g/mL). Since the density is 1.00 g/mL, the mass of 400 mL of solution (200 mL + 200 mL) is 400 g.

Substituting into the equation above, the heat gained by the solution can be represented as:

 $q_{\rm soln} = (4.00 \times 10^2 \text{ g})(4.184 \text{ J/g} \cdot ^{\circ}\text{C})(T_{\rm f} - 20.48 ^{\circ}\text{C})$

Next, let's calculate q_{rxn} , the heat released when 200 mL of 0.862 *M* HCl are mixed with 200 mL of 0.431 *M* Ba(OH)₂. The equation for the neutralization is:

 $2\text{HCl}(aq) + \text{Ba}(\text{OH})_2(aq) \longrightarrow 2\text{H}_2\text{O}(l) + \text{Ba}\text{Cl}_2(aq)$

There is exactly enough $Ba(OH)_2$ to neutralize all the HCl. Note that 2 mole HCl is stoichiometrically equivalent to 1 mole $Ba(OH)_2$, and that the concentration of HCl is double the concentration of $Ba(OH)_2$. The number of moles of HCl is:

$$(2.00 \times 10^2 \text{ mL}) \times \frac{0.862 \text{ mol HCl}}{1000 \text{ mL}} = 0.172 \text{ mol HCl}$$

The amount of heat released when 1 mole of H^+ is neutralized is given in the problem (-56.2 kJ/mol). The amount of heat liberated when 0.172 mole of H^+ is neutralized is:

$$q_{\rm rxn} = 0.172 \text{ mol} \times \frac{-56.2 \times 10^3 \text{ J}}{1 \text{ mol}} = -9.67 \times 10^3 \text{ J}$$

Finally, knowing that the heat lost by the reaction equals the heat gained by the solution, we can solve for the final temperature of the mixed solution.

$$q_{\text{soln}} = -q_{\text{rxn}}$$

$$(4.00 \times 10^2 \text{ g})(4.184 \text{ J/g} \cdot ^\circ\text{C})(T_{\text{f}} - 20.48^\circ\text{C}) = -(-9.67 \times 10^3 \text{ J})$$

$$(1.67 \times 10^3)T_{\text{f}} - (3.43 \times 10^4) = 9.67 \times 10^3 \text{ J}$$

$$T_{\text{f}} = 26.3^\circ\text{C}$$

{Alternate strategy: the net ionic equation for the $HCl + Ba(OH)_2$ reaction is:

$$2\mathrm{H}^{+}(aq) + 2\mathrm{OH}^{-}(aq) \longrightarrow 2\mathrm{H}_{2}\mathrm{O}(l)$$

before dividing by 2--i.e. the reaction

$$2\text{HCl}(aq) + \text{Ba}(\text{OH})_2(aq) \longrightarrow 2\text{H}_2\text{O}(l) + \text{Ba}\text{Cl}_2(aq)$$

creates $2H_2O$, so its ΔH° is 2x(-56.2 kJ/mol) = -112.4 kJ/mol. You can then go from there to the final answer.}

5.45 Strategy: Our goal is to calculate the enthalpy change for the formation of monoclinic sulfur from rhombic sulfur. To do so, we must arrange the equations that are given in the problem in such a way that they will sum to the desired overall equation. This requires reversing the second equation and changing the sign of its ΔH° value.

Solution:

	Rea	<u>ction</u>	ΔH° (kJ/mol)	
	S(rh	nombic) + $O_2(g) \rightarrow SO_2(g)$	-296.06	
	$SO_2(g) \rightarrow S(monoclinic) + O_2(g)$		296.36	
	S(rh	nombic) \rightarrow S(monoclinic)	$\Delta H_{\rm rxn}^{\circ}$ – v.30 kJ/mol	
5.62	(a)	$\Delta H^{\circ} = \left[2\Delta H_{\rm f}^{\circ}(120) + 2\Delta H_{\rm f}^{\circ}(1120) \right] -$	$[\mathcal{L}_{f}^{\circ}(\mathcal{L}_{2}^{1}\mathcal{L}_{4}^{\circ}) + \mathcal{L}_{f}^{\circ}(\mathcal{L}_{2}^{\circ})]$	
		$\Delta H^\circ = [(2)(-393.5 \text{ kJ/mol}) + (2)(-285.8 \text{ kJ/mol})]$	8 kJ/mol)] – [(1)(52.3 kJ/mol) + (3)(0)]	
		$\Delta H^{\circ} = -1411 \text{ kJ/mol}$		
	(b)	$\Delta H^{\circ} = \left[2\Delta H_{\rm f}^{\circ}(112 \odot) + 2\Delta H_{\rm f}^{\circ}(3 \odot 2) \right] -$	$[2221]_{f}^{\circ}(1123) + 3221]_{f}^{\circ}(123)$	
		$\Delta H^\circ = [(2)(-285.8 \text{ kJ/mol}) + (2)(-296.1)]$	kJ/mol] - [(2)(-20.15 kJ/mol) + (3)(0)]	
		$\Delta H^{\circ} = -1124 \text{ kJ/mol}$		
6.20	(a)	$\lambda = \frac{c}{v}$		
		$\lambda = \frac{hc}{E} = \frac{3.00 \times 10^8 \text{ m/s}}{9.87 \times 10^{15} \text{ s}^{-1}} = 3$	$04 \times 10^{-8} \text{ m} = 30.4 \text{ nm}$	

- (b) Checking Figure 6.1 of the text, you should find that the visible region of the spectrum runs from 400 to 700 nm. 30.4 nm is in the **ultraviolet** region of the spectrum.
- (c) E = hv. Substitute the frequency (v) into this equation to solve for the energy of one quantum associated with this frequency.

$$\boldsymbol{E} = h\nu = (6.63 \times 10^{-34} \text{ J} \cdot \text{s})(9.87 \times 10^{15} \text{ s}^{-1}) = 6.54 \times 10^{-18} \text{ J}$$

6.34 In this problem $n_i = 5$ and $n_f = 2$. Using Equation 6.6,

$$\Delta E = -2.18 \times 10^{-18} \text{ J} \left(\frac{1}{2^2} - \frac{1}{5^2}\right) = -4.58 \times 10^{-19} \text{ J}$$

The sign of ΔE means that this is energy associated with an emission process.

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{4.58 \times 10^{-19} \text{ J}} = 4.34 \times 10^{-7} \text{ m} = 434 \text{ nm}$$