

2017 SSSS Chemistry Lab Solutions Manual

Acton Boxborough Regional High School, MA



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The physical properties and stoichiometry problems were written by me, Allen Wang, 19wangal@abschools.org. Since Chris and Antonio were unable to edit the final copy of this test, I take the blame for all the spelling and grammar errors.

Problem solving in chemistry may be difficult at times: some problems may not crack immediately and many require a deep understanding of the fundamental concepts. Don't be discouraged by the test if it is hard. *Learn* to solve the problems and understand the motivation and reasoning behind them. If you have any questions about this test or are interested in test trading during the school year send Antonio or me (Allen) an email.

Enjoy!

General Principles of Chemistry and Physical Properties Multiple Choice

1. C. Br₂
2. D. CuSO₄(H₂O)_{5(s)}
3. B. flame color
4. A. Al
5. B. NH₄OH_(aq)
6. B. heat capacity
7. A. ohm
8. A. 55.6 kJ/K
9. D. the current through a conductor between two points is directly proportional to the voltage across the two points.
10. B. a 1 m of copper wire with 1 mm diameter
11. B. AgNO₃ + NaCl
12. C. also known as the elastic modulus
13. C. sp³
14. D. diamagnetism
15. C. I and III only

General Principles of Chemistry and Physical Properties Open Response and Calculations

- Solution:** Since the molecules get larger as the second atom moves down a group, we would expect melting points to increase down a group.
 - (No partial credit) Due to strong hydrogen bonding in H_2O , the order is $\boxed{\text{H}_2\text{S}, \text{H}_2\text{Se}, \text{H}_2\text{O}}$.
 - (No partial credit) No hydrogen bonding involved so we expect, CO_2 , SiO_2 , GeO_2 , but this is incorrect. At STP, CO_2 is a gas and SiO_2 is sand/quartz which is a network covalent solid. From the strong network bonds, SiO_2 has a higher melting point than GeO_2 . Therefore the correct order is: $\boxed{\text{CO}_2, \text{GeO}_2, \text{SiO}_2}$.
 - (No partial credit) This question involves a bit of trivia. At STP, bromine is a liquid, helium is a gas, while lithium, aluminum, and gallium are solids. Gallium melts at (human) body temperature and thus has the lowest melting point of the solids. Aluminum is a metal used in airplanes and aluminum composite materials are used on spacecrafts so it has the highest melting point of the solids. Thus, the correct order is $\boxed{\text{He}, \text{Br}_2, \text{Ga}, \text{Li}, \text{Al}}$.
- Solution:** Please tell me you have these memorized or written down somewhere.
 - (No partial credit) Electrical conductivity: the degree to which a specified material conducts electricity, calculated as the ratio of the current density in the material to the electric field that causes the flow of current. It is the reciprocal of the resistivity. The SI unit is siemens per meter.
 - (No partial credit) Electrical resistance: a measure of the difficulty to pass an electric current through that conductor. The inverse quantity is electrical conductance. The SI unit is ohms.
 - (No partial credit) Brittleness: the tendency of a material to fracture without significant plastic deformation.
 - (No partial credit) Paramagnetism: the attraction between a material and a magnetic field.
 - (No partial credit) Magnetism: a physical phenomenon produced by the motion of electric charge, resulting in attractive and repulsive forces between objects.
 - (No partial credit) Diamagnetism: the repulsion between a material and a magnetic field.
 - (No partial credit) Color: the property possessed by an object of producing different sensations on the eye as a result of the way the object reflects or emits light.
- Solution:** Recall that $\text{Ca}(\text{OH})_2$ is only slightly soluble (this was used in one of the multiple choice questions). Using K_{sp} requires all of the information to be converted to molarity, M .
 - Since the conditions are at STP, and the assumption is applying, the molarity of OH^- is equal to the molality in 1.0 L of water. So, $[\text{OH}^-] = 0.022 \text{ M}$ and $[\text{Ca}^{2+}] = 0.011 \text{ M}$. Therefore,

$$K_{\text{sp}} = [\text{Ca}^{2+}][\text{OH}^-]^2 = 0.011 \cdot 0.022^2 = \boxed{5.5 \cdot 10^{-6}}$$

- (b) The solution is saturated, so we can use K_{sp} to calculate the pH. The concentration of hydroxide ions is therefore the same as in part a. So

$$\text{pH} = 14 + \log([\text{OH}^-]) = 14 + \log(0.0222) = \boxed{12.346.}$$

- (c) The solution is basic so phenolphthalein will turn pink.

4. **Solution:** This problem involves a little bit of physics. Recall that $F = ma$ where F is the force in N, m is the mass in kg, and a is the acceleration in m/s^2 . This is the type of knowledge all scioly people should know.

- (a) (2 points for physics portion, 2 points for answer) Using the formula $F = ma$, since the prism is falling on Earth, $a = 9.8 \text{ m/s}^2$, the acceleration of gravity. Solving the equation shows that the mass is 162 g. So the density is:

$$\rho = \frac{162 \text{ g}}{3.00 \text{ cm} \cdot 4.00 \text{ cm} \cdot 5.00 \text{ cm}} = \boxed{2.70 \text{ g/cm}^3.}$$

- (b) (1 point for attempt, 2 points for completing first calculation, 2 points for answer) Although this is a typical material science question, pure logic and geometry does the trick here anyways. We can calculate the atomic density of aluminum, multiply by a factor of Avogadro's number, and use the molar mass. The atomic density is

$$\frac{\text{number of atoms per cell}}{\text{volume of the cell}} = \frac{4 \text{ atoms}}{(404 \text{ pm})^3} \cdot \left(\frac{10^{10} \text{ pm}}{1 \text{ cm}}\right)^3 = 6.07 \cdot 10^{22} \text{ atoms/cm}^3.$$

We multiply this by the atomic mass, 26.98 g/mol, and divide by $6.022 \cdot 10^{23}$ atoms/mol, to get the final density of

$$\frac{6.07 \cdot 10^{22} \text{ atoms}}{\text{cm}^3} \cdot \frac{26.98 \text{ g}}{\text{mol}} \cdot \frac{\text{mol}}{6.022 \cdot 10^{23} \text{ atoms}} = \boxed{2.72 \text{ g/cm}^3,}$$

which is fairly close to our answer in part a.

- (c) (No partial credit) Like national tests, this test includes funky units. With the unit conversion values memorized or on the cheatsheet, this is an easy job for the calculator:

$$\frac{2.72 \text{ g}}{\text{cm}^3} \cdot \left(\frac{2.54 \text{ cm}}{\text{in}}\right)^3 \cdot \frac{0.00220 \text{ lb}}{\text{g}} \approx \boxed{0.0981 \text{ lb/in}^3.}$$

5. **Solution:** (1 point for formula, 1 point for correct R with units given, 2 points for calculation, 5 points for answer) Recall the Clausius-Clapeyron equation:

$$\Delta \ln(P) = \frac{-\Delta H_{\text{vap}}}{R} \left(\Delta \frac{1}{T}\right).$$

Since the normal boiling point of X occurs when vapor pressure equals 760 torr, T is 313 K, and $R = 8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}}$,

$$\ln\left(\frac{760}{213}\right) = -\frac{31400 \text{ kJ/mol}}{8.314 \text{ J/(mol} \cdot \text{K)}} \left(\frac{1}{T} - \frac{1}{313 \text{ K}}\right).$$

Solving the equation shows that $T = \boxed{350. \text{ K}}$

6. **Solution:** Note that CaCl_2 is ionic, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ is table sugar, Zr and Cu are metals, Si is a semiconductor, and NH_3 is a gas that acts as a weak base when dissolved in water.

- (a) (No partial credit) Only NH_3 exists as a gas at STP.
- (b) (No partial credit) Either one of Zr or Cu can be first or second, Si is third, and either one of CaCl_2 or $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ can be fourth or fifth.
- (c) (No partial credit) Reverse order of part b.
- (d) (No partial credit) The correct order is NH_3 , CaCl_2 , $\text{C}_{12}\text{H}_{22}\text{O}_{11}$.
- (e) (No partial credit) Since CaCl_2 has a Van't Hoff factor of 3, it changes the boiling point of water the most.
- (f) (1 point for formula, 1 point for recognizing Van't Hoff factor, 1 point for answer) Recall the formula $\Delta T = K_f \cdot m \cdot i$, where ΔT is the freezing point depression, K_f is the freezing point depression constant, m is the molality of the solution, and i is the Van't Hoff factor. Plugging in the data given shows that:

$$\Delta T = \frac{1.86 \text{ }^\circ\text{C}}{m} \cdot \frac{25 \text{ g CaCl}_2}{0.975 \text{ kg}} \cdot \frac{\text{mol CaCl}_2}{110.98 \text{ g}} \cdot 3 = 1.3 \text{ }^\circ\text{C}.$$

Thus, the new freezing point will be $\boxed{-1.3 \text{ }^\circ\text{C}}$.

- (g) (No partial credit) The answer is NH_3 . The reaction is $\text{NH}_{3(\text{g})} + \text{H}_2\text{O}_{(\ell)} \rightarrow \text{NH}_{4(\text{aq})}^+ + \text{H}_3\text{O}_{(\text{aq})}^+$.

7. **Solution:**

- (a) Na_2CO_3
- (b) $\text{Ca}(\text{NO}_2)_2$
- (c) H_2SO_3
- (d) Hg_2O
- (e) $\text{Pb}(\text{Cr}_2\text{O}_7)_2$
- (f) $\text{CuSO}_4(\text{H}_2\text{O})_5$
- (g) C_3H_6
- (h) P_4O_{10}
- (i) Silver Nitrate
- (j) Iron(II) Bromide
- (k) Ethane
- (l) Acetic Acid or Ethanoic Acid

- (m) Sodium Dihydrogen Phosphate
- (n) Ammonium Permanganate
- (o) Cobalt(II) Thiosulfate
- (p) Sulfur Trioxide
8. **Solution:** The calculations should be self-explanatory. Significant figures are used only at the end.
- (a) (1 point awarded if only coefficients are wrong, 3 points otherwise) $2\text{Cr}_2\text{O}_7^{2-} + 3\text{C}_2\text{H}_5\text{OH} + 16\text{H}^+ \rightarrow 4\text{Cr}^{3+} + 3\text{CH}_3\text{COOH} + 11\text{H}_2\text{O}$
- (b) (1 point awarded if one of the two colors are correct, otherwise 3 points) The color changes from orange to green.
- (c) (No partial credit) $0.750 \text{ g} \cdot \frac{1 \text{ mol}}{294.2 \text{ g}} \cdot \frac{1}{0.250 \text{ L}} = \boxed{0.010 \text{ M}}$
- (d) (2 points awarded per calculation, up to 6 points total) $0.02025 \text{ L} \cdot \frac{0.0102 \text{ mol}}{\text{L}} = 2.066 \cdot 10^{-4} \text{ mol Cr}_2\text{O}_7^{2-}$ total.

$$2.066 \cdot 10^{-4} \text{ mol Cr}_2\text{O}_7^{2-} \cdot \frac{[3 \text{ mol C}_2\text{H}_5\text{OH}]}{[2(\text{Cr}_2\text{O}_7^{2-})]} = 3.099 \cdot 10^{-4} \text{ mol C}_2\text{H}_5\text{OH in 10 mL.}$$

$$\text{Finally, } 3.099 \cdot 10^{-4} \text{ mol C}_2\text{H}_5\text{OH} \cdot \frac{100}{10} = \boxed{3.10 \cdot 10^{-3} \text{ mol C}_2\text{H}_5\text{OH in 100 mL.}}$$

- (e) (No partial credit) $0.600 \text{ mL} \cdot \frac{0.966 \text{ g}}{\text{cm}^3} = 0.5796 \text{ g}$. Then, $3.099 \cdot 10^{-3} \text{ mol} \cdot \frac{46.7 \text{ g}}{\text{mol}} = 0.1428 \text{ g C}_2\text{H}_5\text{OH}$
- $\Rightarrow \boxed{24.6\% \text{ ethanol.}}$

Thermochemistry Multiple Choice

1. D. $C_4H_{10(g)}$
2. B. intramolecular forces
3. C. convection
4. B. radiation
5. A. joule
6. C. freezing point
7. B. is susceptible to decomposition
8. D. deposition
9. B. enthalpy
10. D. temperature
11. A. a liquid that will boil when altitude increases
12. A. exist at the same time
13. C. $41.6\text{ }^\circ\text{C}$
14. A. 61.13 kJ
15. C. I and II only

Thermochemistry Open Response and Calculations

1. **Solution:** (1 points awarded for ideas in the first sentence, 3 more points if enthalpy and heat content are mentioned) Cousteau's answer is scientifically inaccurate because he describes the Antarctic as a source of cold. Cold is not a quantity that can be released from any source; rather, cold objects are simply lacking in enthalpy, or heat content. His explanation should state that warmer water decreases in temperature by transferring its heat to the cooler Antarctic water.

2. **Solution:**

(a) (half credit given up to the gram to ounces calculation, rest of the credit given for the correct answer) To find the heat that the water absorbed, we have:

$$q = mC\Delta T.$$

$$q = 260. \text{ g} \times \frac{1 \text{ cal}}{\text{g}^\circ\text{C}} \times (85.9 \text{ }^\circ\text{C} - 20.0 \text{ }^\circ\text{C}) = 17100 \text{ cal}.$$

Remember that a Calorie with a capital C is actually equivalent to a kilocalorie, or 1000 calories. Converting calories to Calories, we have:

$$17100 \text{ cal} \times \frac{1 \text{ Cal}}{1000 \text{ cal}} = 17.1 \text{ Cal}.$$

Converting grams to ounces, we have:

$$3.50 \text{ g} \times \frac{1 \text{ ounce}}{28.35 \text{ g}} = 0.123 \text{ ounces}.$$

The ratio of Calories to ounces is:

$$\frac{17.1 \text{ Cal}}{0.123 \text{ ounces}} = \boxed{139 \text{ Calories/ounce.}}$$

Clearly, this is not 125 Calories per ounce as stated by the company. Therefore, the company is incorrect.

(b) (1 point for attempt to write equation, 3 points for correct answer) Let m be the number of grams of fat in the 3.50 g sample. From part (a), we know that the sample contains 17.1 Calories. First converting to kilojoules, we have:

$$17.1 \text{ Cal} \times \frac{4.184 \text{ kJ}}{\text{Cal}} = 71.5 \text{ kJ}.$$

We set up an equation to calculate the amount of fat. The mass that is not composed of fats, or $3.50 \text{ g} - m$, must be carbohydrates or proteins, both of which contain 16.7 kilojoules per gram. Thus, we have:

$$16.7 \text{ kJ/g} \times (3.50 \text{ g} - m) + 37.7m \text{ kJ/g} = 71.5 \text{ kJ}.$$

$$58.5 \text{ kJ} - 16.7m \text{ kJ/g} + 37.7m \text{ kJ/g} = 71.5 \text{ kJ}.$$

$$21.0m \text{ kJ/g} = 13.0 \text{ kJ}.$$

$$m = \frac{13.0}{21.0} \text{ g} = 0.619 \text{ g}.$$

This is for our 3.50 gram sample. To find this per ounce, we find the amount per gram and then convert to ounces:

$$0.619 \text{ g} \times \frac{1}{3.50} \times \frac{28.35 \text{ g}}{1 \text{ ounce}} = \boxed{5.01 \text{ grams/ounce.}}$$

Here, the company is telling the truth.

There are a variety of ways to solve this problem, so there may be some round-off error. Regardless, the difference from 5.00 is too minimal to say that the company is not telling the truth, especially since the value was likely rounded to a whole number.

- (c) (1 point for talking about macromolecules, 3 more points for full response) The energy was initially located within the bonds of the carbohydrates, proteins, and fats. The energy used to initiate the combustion broke the bonds in these macromolecules, converting them to heat. A characteristic that allows certain substances to have a higher heat concentration would be a higher number of bonds, allowing them to store more energy. For example, fats are densely packed with C-H bonds, allowing them to store great amounts of energy.
- (d) (1 point for mentioning coffee cup, 2 more points for full answer) One possible source of error is that the coffee cup calorimeter absorbed some of the heat produced by the combustion. A way to improve the experiment would be to use a calorimeter with a higher insulation than a styrofoam cup. Another possible source of error is that some of the heat escaped from the water into the surroundings, as a coffee cup calorimeter is open at the top. A way to fix this issue would be to instead use a closed calorimeter, such as a bomb calorimeter.

Both of these errors mean that the true number of calories is actually higher than 139 because not all the heat from the combustion was absorbed by the water.

3. **Solution:** Note that here the values for standard heats of formation and standard entropies are calculated with significant figures according to their digits. If they are treated as exact values, more significant figures will be present in the final answer. Exact answers are provided in parentheses with four digits, but other acceptable answers may have more than four.

- (a) (1 point for balanced equation, 2 more points for correct answer) The balanced chemical equation is:



Using the standard heats of formation table, we can find ΔH° for the reaction:

$$\Delta H^\circ = \sum H_{\text{products}}^\circ - \sum H_{\text{reactants}}^\circ$$

$$\Delta H^\circ = -134 \text{ kJ/mol} - (1.88 \text{ kJ/mol} + 2 \times 0 \text{ kJ/mol}) = -136 \text{ kJ/mol}$$

We also need to find the change in entropy of the system. Using the table of standard entropies, we can calculate ΔS° (remember that \sum means the sum of all values):

$$\Delta S^\circ = \sum S_{\text{products}}^\circ - \sum S_{\text{reactants}}^\circ$$

$$\Delta S^\circ = 309.65 \frac{\text{J}}{\text{mol} \times \text{K}} - (2.439 \frac{\text{J}}{\text{mol} \times \text{K}} + 2 \times 223.0 \frac{\text{J}}{\text{mol} \times \text{K}})$$

$$\Delta S^\circ = 309.65 \frac{\text{J}}{\text{mol} \times \text{K}} - 448.4 \frac{\text{J}}{\text{mol} \times \text{K}} = -138.8 \frac{\text{J}}{\text{mol} \times \text{K}}.$$

Converting ΔS° to kilojoules, we get:

$$\Delta S^\circ = -138.8 \frac{\text{J}}{\text{mol} \times \text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = -0.1388 \frac{\text{kJ}}{\text{mol} \times \text{K}}.$$

We convert 25 °C to Kelvin:

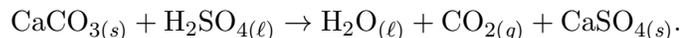
$$T = 25 \text{ }^\circ\text{C} + 273 = 298\text{K}.$$

Now, we can calculate ΔG° :

$$\begin{aligned} \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ. \\ &= -136 \text{ kJ/mol} - 298\text{K} \times -0.1388 \frac{\text{kJ}}{\text{mol} \times \text{K}}. \\ &= -136 \text{ kJ/mol} + 41.36 \text{ kJ/mol}. \\ &= \boxed{-95 \text{ kJ/mol} (-94.53 \text{ kJ/mol exactly}).} \end{aligned}$$

Since $\Delta H^\circ < 0$ and $\Delta G^\circ < 0$, the reaction is exothermic and spontaneous.

- (b) (1 point for balanced equation, 1 point for formula for Gibb's free energy, 2 points for final answer) Our balanced equation is:



First, we calculate ΔH using the values in the standard heats of formation table:

$$\Delta H^\circ = -1433 \text{ kJ/mol} - 394 \text{ kJ/mol} - 286 \text{ kJ/mol} - (-813.8 \text{ kJ/mol} - 1207 \text{ kJ/mol}).$$

$$\Delta H^\circ = -92 \text{ kJ/mol}.$$

We have that 25 °C = 298K from part (a). We then calculate ΔS° :

$$\begin{aligned} \Delta S^\circ &= 106.5 \frac{\text{J}}{\text{mol} \times \text{K}} + 213.6 \frac{\text{J}}{\text{mol} \times \text{K}} + 69.94 \frac{\text{J}}{\text{mol} \times \text{K}} \\ &\quad - \left(88.7 \frac{\text{J}}{\text{mol} \times \text{K}} + 156.9 \frac{\text{J}}{\text{mol} \times \text{K}} \right) \\ &= 144.4 \frac{\text{J}}{\text{mol} \times \text{K}}. \end{aligned}$$

Converting that to kilojoules, we have:

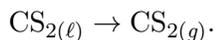
$$\Delta S^\circ = 144.4 \frac{\text{J}}{\text{mol} \times \text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.1444 \frac{\text{kJ}}{\text{mol} \times \text{K}}.$$

Now, we can calculate ΔG° :

$$\begin{aligned} \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ. \\ &= -92 \text{ kJ/mol} - 298\text{K} \times 0.1444 \frac{\text{kJ}}{\text{mol} \times \text{K}}. \\ &= -92 \text{ kJ/mol} - 43.03 \frac{\text{kJ}}{\text{mol} \times \text{K}}. \\ &= \boxed{-135 \text{ kJ/mol} (-135.0 \text{ kJ/mol exactly}).} \end{aligned}$$

Since $\Delta H^\circ < 0$ and $\Delta G^\circ < 0$, the reaction is exothermic and spontaneous.

(c) (No partial credit) Our balanced equation is:



We first determine ΔH° :

$$\Delta H^\circ = +117 \text{ kJ/mol} - 89.5 \text{ kJ/mol} = +28 \text{ kJ/mol}.$$

We have that $25^\circ\text{C} = 298\text{K}$ from the part (a). We now find ΔS° :

$$\Delta S^\circ = 237.98 \frac{\text{J}}{\text{mol} \times \text{K}} - 151.0 \frac{\text{J}}{\text{mol} \times \text{K}} = 87.0 \frac{\text{J}}{\text{mol} \times \text{K}}.$$

Converting that to kilojoules, we have:

$$\Delta S^\circ = 87.0 \frac{\text{J}}{\text{mol} \times \text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.0870 \frac{\text{kJ}}{\text{mol} \times \text{K}}.$$

Now, we can calculate ΔG° :

$$\begin{aligned} \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ \\ &= 28 \text{ kJ/mol} - 298\text{K} \times 0.0870 \frac{\text{J}}{\text{mol} \times \text{K}} \\ &= 28 \text{ kJ/mol} - 25.9 \text{ kJ/mol} \\ &= \boxed{2 \text{ kJ/mol (1.580 kJ/mol exactly)}}. \end{aligned}$$

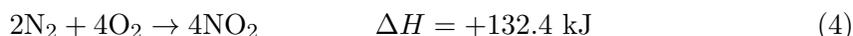
Since $\Delta H^\circ > 0$ and $\Delta G^\circ > 0$, the reaction is endothermic and nonspontaneous.

4. Solution:

(a) (1 point for attempt to eliminate intermediates, 3 more points for correct answer) Our equations are:



If we reverse equation (2) and then multiply it by 4, we get:



Adding this to equation (1) yields:



Multiplying equation (3) by 2 yields the following:



Adding (5) and (6) together yields:



This equation simplifies to our desired equation:



Therefore, for the overall formation, process, $\Delta H = \boxed{+100.6 \text{ kJ}}$.

- (b) (1 point for attempt to do stoic, 2 more points for correct answer) First, we convert 81.0 J to kilojoules:

$$81.0 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.0810 \text{ kJ}.$$

For every 100.6 kJ absorbed, 4 moles of nitric acid are formed. Let x be the moles of nitric acid formed. We have:

$$\frac{0.0810 \text{ kJ}}{100.6 \text{ kJ}} = \frac{x}{4 \text{ mol}}$$
$$x = \frac{0.0810 \times 4}{100.6} \text{ mol} = 0.00322 \text{ mol}.$$

The molar mass of nitric acid is:

$$MM = 1.01 \text{ g/mol} + 14.01 \text{ g/mol} + 3 \times 16.00 \text{ g/mol} = 63.02 \text{ g/mol}.$$

Therefore, the number of grams of nitric acid formed is:

$$0.00322 \text{ mol} \times \frac{63.02 \text{ g}}{1 \text{ mol}} = \boxed{0.203 \text{ g}}.$$

5. Solution:

- (a) (1 point for calculating heat evolved in some steps, 2 points for calculating total heat involved, 2 points for correct answer) We calculate the heat absorbed by the water using the formula $q = mc\Delta T$:

$$q = 156.0 \text{ g} \times \frac{4.184 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times (17.962 \text{ }^\circ\text{C} - 21.945 \text{ }^\circ\text{C}).$$

$$q = 156.0 \text{ g} \times \frac{4.184 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times -3.983 \text{ }^\circ\text{C}.$$

$$q = -2306 \text{ J}.$$

Since the heat absorbed by the benzene is equivalent to the heat released by the water, $q_{\text{rxn}} = -q_{\text{H}_2\text{O}}$. We now find the number of moles of benzene. The molar mass is:

$$MM = 6 \times 12.01 \text{ g/mol} + 6 \times 1.01 \text{ g/mol} = 72.06 \text{ g/mol} + 6.06 \text{ g/mol} = 78.12 \text{ g/mol}.$$

Converting grams of benzene to moles, we have:

$$12.80 \text{ g} \times \frac{1 \text{ mol}}{78.12 \text{ g}} = 0.1639 \text{ mol}.$$

Let T be the melting point of benzene. Let q_1 be the enthalpy change of the benzene when temperature of the benzene rises from $-15.810 \text{ }^\circ\text{C}$ to T . Let q_2 be the enthalpy change of fusion. Let q_3 be the enthalpy change when the temperature of the benzene rises from T to $17.962 \text{ }^\circ\text{C}$. Using the formula $q = mc\Delta T$ and the given molar heat capacities, we can find expressions for q_1 and q_3 in terms of T :

$$q_1 = 0.1639 \text{ mol} \times \frac{118.4 \text{ J}}{\text{mol} \cdot ^\circ\text{C}} \times (T + 15.810 \text{ }^\circ\text{C}) = 19.41T \text{ J/}^\circ\text{C} + 306.8 \text{ J}.$$

$$q_3 = 0.1639 \text{ mol} \times \frac{134.8 \text{ J}}{\text{mol} \cdot ^\circ\text{C}} \times (17.962 \text{ }^\circ\text{C} - T) = 396.8 \text{ J} - 22.09T \text{ J/}^\circ\text{C}.$$

We can find q_2 with the given ΔH_{fus} :

$$q_2 = n\Delta H_{\text{fus}} = 0.1639 \text{ mol} \times \frac{9.866 \text{ kJ}}{1 \text{ mol}} = 1.617 \text{ kJ}.$$

Converting q_2 to joules, we have:

$$q_2 = 1.617 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 1617 \text{ J}.$$

Adding all the steps together, we have:

$$q_{\text{rxn}} = q_1 + q_2 + q_3.$$

$$2306 \text{ J} = 19.41 \text{ J/}^\circ\text{C} + 306.8 \text{ J} + 1617 \text{ J} + 396.8 \text{ J} - 22.09 \text{ J/}^\circ\text{C}.$$

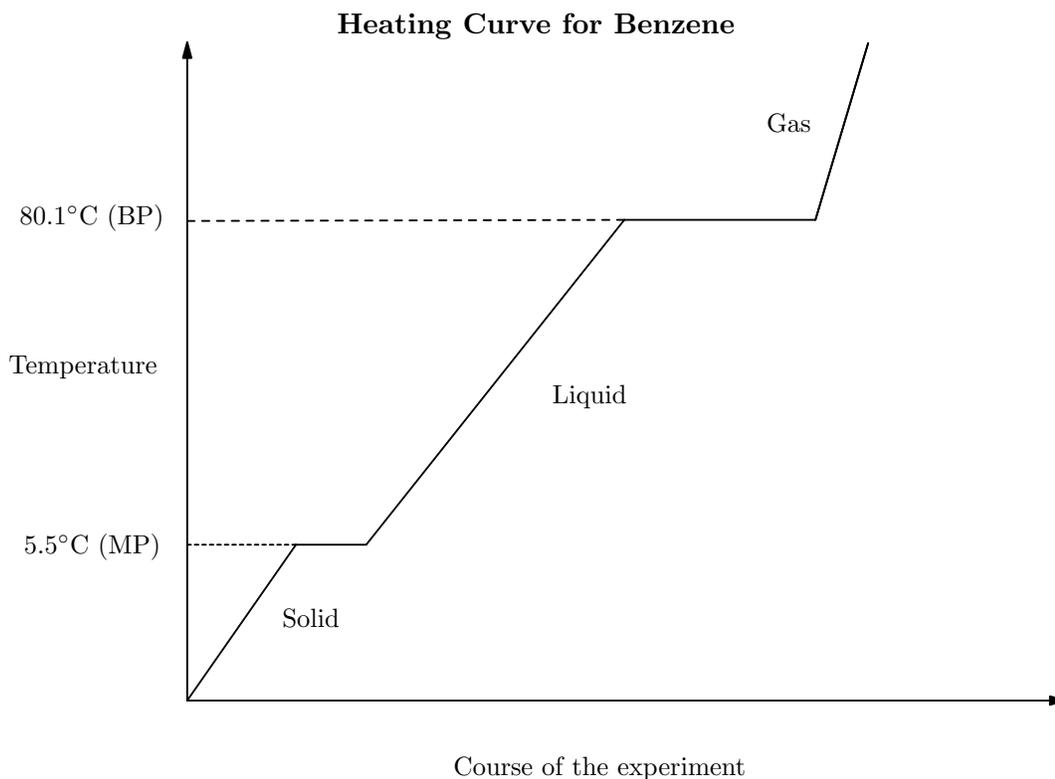
$$2306 \text{ J} = -2.68T \text{ J/}^\circ\text{C} + 2321 \text{ J}.$$

$$-15 \text{ J} = -2.68T \text{ J/}^\circ\text{C}.$$

$$T = \frac{-15}{-2.68} = \boxed{5.6 \text{ }^\circ\text{C}}.$$

A quick note about this problem: if you are one of those people that refuses to use correct significant figures, you will get this question wrong by quite a bit. A single rounding error will mess up you significantly. This problem took about an hour to get right. Notice how the numbers in the problem have four or five significant figures and the answer comes out with two?

- (b) (1 point for every two items included, 5 points total for complete, correct answer) The graph:



Lovely. Please note that this diagram is by no means to scale. Persons attempting to find scale in this diagram will be prosecuted. Persons attempting to find motive will be banished. Persons attempting to find plot (what a pun) will be shot.

(Paragraph in reference to the disclaimer in *Huckleberry Finn* by Mark Twain.)

- (c) (1 point for each of q_1, q_2, q_3, q_4 , 1 point for final answer) We first find the number of moles of benzene:

$$41.05 \text{ g} \times \frac{1 \text{ mol}}{78.11 \text{ g}} = 0.5255 \text{ mol.}$$

Note that in the following equations, n represents the number of moles, $C_{p(s)}$ represents the molar heat capacity of the solid phase, and $C_{p(l)}$ represents the molar heat capacity of the liquid phase.

Let q_1 be the change in enthalpy of benzene from -15.810°C to 5.6°C . We have:

$$q_1 = nC_{p(s)}\Delta T.$$

$$q_1 = 0.5255 \text{ mol} \times 118.4 \frac{\text{J}}{\text{mol}^\circ\text{C}} \times (5.6^\circ\text{C} + 15.810^\circ\text{C}).$$

$$q_1 = 0.5255 \text{ mol} \times 118.4 \frac{\text{J}}{\text{mol}^\circ\text{C}} \times 21.4^\circ\text{C} = 1330 \text{ J.}$$

Let q_2 be the change in enthalpy during fusion.:

$$q_2 = n\Delta H_{\text{fus}} = 0.5255 \text{ mol} \times 9.866 \text{ kJ/mol} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 5185 \text{ J.}$$

Let q_3 be the change in enthalpy from 5.6°C to 80.1°C . We have:

$$q_3 = nC_{p(l)}\Delta T.$$

$$q_3 = 0.5255 \text{ mol} \times 134.8 \frac{\text{J}}{\text{mol}^\circ\text{C}} \times (80.1^\circ\text{C} - 5.6^\circ\text{C}).$$

$$q_3 = 0.5255 \text{ mol} \times 134.8 \frac{\text{J}}{\text{mol}^\circ\text{C}} \times 74.5^\circ\text{C} = 5280 \text{ J.}$$

Let q_4 be the change in enthalpy during boiling. We have:

$$q_4 = n\Delta H_{\text{boil}} = 0.5255 \text{ mol} \times 30.77 \text{ kJ/mol} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 16170 \text{ J.}$$

We now add these all together to find the total amount of heat absorbed by the benzene:

$$q_{\text{C}_6\text{H}_6} = q_1 + q_2 + q_3 + q_4.$$

$$q_{\text{C}_6\text{H}_6} = 1330 \text{ J} + 5185 \text{ J} + 5280 \text{ J} + 16170 \text{ J} = 27970 \text{ J.}$$

Now let us find the heat released by the water. The number of grams of water is:

$$27.18 \text{ mol} \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 489.2 \text{ g.}$$

For the benzene to absorb heat from the water, the water must always be at a higher temperature than the benzene. This means that the minimum possible temperature

of the water by the end is 80.1 °C, the boiling point of benzene. Let T be the initial temperature of the water. We can find the heat released by the water in terms of T :

$$q_{\text{H}_2\text{O}} = mc\Delta T.$$

$$q_{\text{H}_2\text{O}} = 489.2 \text{ g} \times \frac{4.184 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times (80.1 \text{ } ^\circ\text{C} - T).$$

$$q_{\text{H}_2\text{O}} = 164000 \text{ J} - 2047T \text{ J}/^\circ\text{C}.$$

Since the benzene absorbed all the heat the water released, $q_{\text{C}_6\text{H}_6} = -q_{\text{H}_2\text{O}}$. Therefore, using our values for $q_{\text{C}_6\text{H}_6}$ and $q_{\text{H}_2\text{O}}$, we can solve for T :

$$27970 \text{ J} = -(164000 \text{ J} - 2047T \text{ J}/^\circ\text{C}).$$

$$27970 \text{ J} = -164000 \text{ J} + 2047T \text{ J}/^\circ\text{C}.$$

$$192000 \text{ J} = 2047T \text{ J}/^\circ\text{C}.$$

$$T = \frac{192000}{2047} \text{ } ^\circ\text{C} = \boxed{93.8 \text{ } ^\circ\text{C}}.$$

6. **Solution:** (1 point for knowing that Gibb's must be equal to 0, 1 point for writing formulas, 3 more points for correct answer) $\text{Br}_{2(\ell)} \rightarrow \text{Br}_{2(g)}$ is spontaneous above the boiling point and nonspontaneous below the boiling point. ΔG is negative for spontaneous reactions and positive for nonspontaneous reactions. The boiling point marks the boundary between spontaneous and nonspontaneous vaporization; thus, $\Delta G = 0$ for $\text{Br}_{2(\ell)} \rightarrow \text{Br}_{2(g)}$ at the boiling point of bromine. Although ΔH and ΔS vary with temperature, we are told we can approximate using ΔH° and ΔS° , and the boiling point of bromine is relatively close to 25 °C. We have:

$$\Delta G = \Delta H - T\Delta S.$$

$$0 = \Delta H^\circ - T\Delta S^\circ.$$

$$T\Delta S^\circ = \Delta H^\circ.$$

$$T = \frac{\Delta H^\circ}{\Delta S^\circ}.$$

Now, we calculate ΔH° :

$$\Delta H^\circ = \sum H^\circ_{\text{products}} - \sum H^\circ_{\text{reactants}}.$$

$$\Delta H^\circ = +30.9 \text{ kJ/mol} - 0 \text{ kJ/mol} = +30.9 \text{ kJ/mol}.$$

Converting, we have:

$$\Delta H^\circ = +30.9 \text{ kJ/mol} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = +30900 \text{ J/mol}.$$

We now find ΔS :

$$\Delta S^\circ = \sum S^\circ_{\text{products}} - \sum S^\circ_{\text{reactants}}.$$

$$\Delta S^\circ = 245.3 \frac{\text{J}}{\text{mol} \times \text{K}} - 152.3 \frac{\text{J}}{\text{mol} \times \text{K}} = 93.0 \frac{\text{J}}{\text{mol} \times \text{K}}.$$

We can now find the temperature in Kelvin:

$$T = \frac{+30900 \text{ J/mol}}{93.0 \frac{\text{J}}{\text{mol}\times\text{K}}} = 332\text{K}.$$

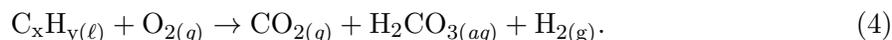
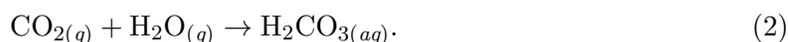
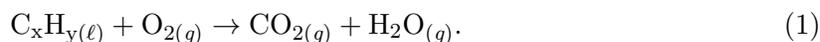
Converting to Celsius, we get:

$$332\text{K} - 273 = \boxed{59^\circ\text{C}}.$$

7. Solution:

- (a) (No partial credit) The blue section is the solid phase, the green section is the liquid phase, and the yellow section is the gas phase.
- (b) (No partial credit) Sublime into a gas.
- (c) (No partial credit) 5.11 atm.
- (d) (No partial credit) Critical point.
- (e) (No partial credit) Increase the pressure or lower the temperature.

8. Solution: (5 points for empirical formula, 5 moe points for correct answer) Our equations are:



The first and fourth ones are unbalanced. We find the number of moles of carbonic acid produced:

$$16.88 \text{ g} \times \frac{1 \text{ mol}}{62.03 \text{ g}} = 0.2721 \text{ mol}.$$

Carbonic acid and carbon dioxide are in a 1 : 1 ratio in equation (2); therefore, there also are 0.2721 moles of CO₂. This is 11.14% of the total carbon dioxide produced by reaction (1). To find the total amount of carbon dioxide, we divide by 0.1114 (11.14%):

$$\frac{0.2721 \text{ mol}}{0.1114} = 2.443 \text{ mol}.$$

For equation (1), there are 2.443 moles of CO₂. As there is one carbon atom per molecule of CO₂, there are 2.443 moles of carbon atoms in the reaction. CO₂ is the only source of carbon in the reaction, so C_xH_y also contains 2.443 moles of carbon atoms. The mass of carbon in the 34.27 g sample is:

$$2.443 \text{ mol} \times \frac{12.01 \text{ g}}{1 \text{ mol}} = 29.34 \text{ g}.$$

Therefore, the mass of hydrogen in the hydrocarbon is:

$$34.27 \text{ g} - 29.34 \text{ g} = 4.93 \text{ g}.$$

Converting to moles of hydrogen atoms, we get:

$$4.93 \text{ g} \times \frac{1 \text{ mol}}{1.01 \text{ g}} = 4.88 \text{ g}.$$

There are 2.443 moles of carbon and 4.88 moles of hydrogen in the sample of C_xH_y . To find the empirical formula, we divide both numbers by 2.443:

$$C : \frac{2.443 \text{ mol}}{2.443 \text{ mol}} = 1. \qquad H : \frac{4.88 \text{ mol}}{2.443 \text{ mol}} = 2.00.$$

There are twice as many moles of hydrogen as there are moles of carbon. Therefore, $y = 2x$, and the empirical formula for the hydrocarbon is CH_2 . For moles of C_xH_y , we have:

$$\frac{34.27 \text{ g}}{(12.01x + 1.01y) \text{ g/mol}} = \frac{34.27 \text{ g}}{(12.01x + 2.02x) \text{ g/mol}} = \frac{34.27}{14.03x} \text{ mol}.$$

By Hess's Law, we can add ΔH° for equations (1), (2) and (3) together to get ΔH° equation (4). We now wish to find the change in heat for each reaction. The enthalpy change for equation (1) is:

$$\Delta H_1 = \frac{-4294 \text{ kJ}}{1 \text{ mol}} \times \frac{34.27}{14.03x} \text{ mol} = -\frac{10490 \text{ kJ}}{x}.$$

For equation (2), we find that the standard heat of formation for the reaction is:

$$\Delta H^\circ = -699.65 \text{ kJ/mol} - (-394 \text{ kJ/mol} - 242 \text{ kJ/mol}) = -63.65 \text{ kJ/mol}.$$

Here the standard heats of formation are being treated as exact values, but they may also be treated as having measured values with significant figures. Now, with the 0.2721 moles of H_2CO_3 in the experiment:

$$\Delta H_2 = \frac{-63.65 \text{ kJ}}{1 \text{ mol}} \times 0.2721 \text{ mol} = -17.33 \text{ kJ}.$$

We know that the standard heat of formation for $H_2O_{(g)} = -242 \text{ kJ/mol}$. We have to reverse the sign when doing the decomposition. The change in enthalpy is for equation (3) is:

$$\Delta H_3 = +242 \text{ kJ/mol} \times (2.443 \text{ mol} - 0.2721 \text{ mol}) = +523.4 \text{ kJ}.$$

The total change in enthalpy is $\Delta H_{rxn} = -990.6 \text{ kJ}$. Thus, we have:

$$\begin{aligned} \Delta H_{rxn} &= \Delta H_1 + \Delta H_2 + \Delta H_3 \\ -990.6 \text{ kJ} &= -\frac{10490 \text{ kJ}}{x} - 17.33 \text{ kJ} + 523.4 \text{ kJ} \\ -1496.7 \text{ kJ} &= -\frac{10490 \text{ kJ}}{x} \\ x &= \frac{-10490}{-1496.7} = 7.009. \end{aligned}$$

Therefore, the formula for the hydrocarbon is $\boxed{C_7H_{14}}$.

9. Solution:

(a) (No partial) We first calculate the number of moles of the sodium sulfide:

$$9.15 \text{ g} \times \frac{1 \text{ mol}}{78.0 \text{ g}} = 0.117 \text{ mol}.$$

We then find the heat released by the sodium sulfide:

$$q_{\text{Na}_2\text{S}} = n\Delta H = 0.117 \text{ mol} \times (-63.5 \text{ kJ/mol}) = -7.43 \text{ kJ}.$$

Converting to joules, we have:

$$q_{\text{Na}_2\text{S}} = -7.43 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = -7430 \text{ J}.$$

We now determine the mass of the water:

$$75.0 \text{ mL} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{1 \text{ g}}{1 \text{ cm}^3} = 75.0 \text{ g}.$$

We find the temperature change of the water as it absorbs heat from the sodium sulfide:

$$-q_{\text{Na}_2\text{S}} = q_{\text{H}_2\text{O}} = 7430 \text{ J}.$$

$$q_{\text{H}_2\text{O}} = mc\Delta T.$$

$$7430 \text{ J} = 75.0 \text{ g} \times \frac{4.184 \text{ J}}{\text{g}^\circ\text{C}} \times \Delta T.$$

$$\Delta T = \frac{7430}{75.0 \times 4.184} ^\circ\text{C} = 23.7 ^\circ\text{C}.$$

Thus, we know that the final temperature is $25.0 ^\circ\text{C} + 23.7 ^\circ\text{C} = \boxed{48.7 ^\circ\text{C}}$.

- (b) (3 points for equilibrium temperature, 3 more points for the final answer) There are 75.0 milliliters of sodium sulfide solution, 90.0 milliliters of hydrochloric acid solution, and 165.0 milliliters of total solution after they are mixed. We are given that the density of water is exactly 1 g/cm^3 , so we can convert of each these volumes to mass:

$$75.0 \text{ mL} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{1 \text{ g}}{1 \text{ cm}^3} = 75.0 \text{ g}.$$

$$90.0 \text{ mL} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{1 \text{ g}}{1 \text{ cm}^3} = 90.0 \text{ g}.$$

$$165.0 \text{ mL} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{1 \text{ g}}{1 \text{ cm}^3} = 165.0 \text{ g}.$$

Ignoring the heat transfer caused by the chemical reaction, we can determine the equilibrium temperature of the system after the hydrochloric acid solution has been added.

Let q be the enthalpy change of the sodium sulfide solution. As the hydrochloric acid solution absorbs heat from the sodium sulfide solution until equilibrium temperature is reached, the enthalpy change of the hydrochloric acid solution is $-q$.

Now, let T be the equilibrium temperature. We can now solve for T :

$$q = 75.0 \text{ g} \times \frac{1 \text{ cal}}{\text{g}^\circ\text{C}} \times (T - 48.7 ^\circ\text{C}).$$

$$-q = 90.0 \text{ g} \times \frac{1 \text{ cal}}{\text{g}^\circ\text{C}} \times (T - 20.0 ^\circ\text{C}).$$

Adding the two equations together, we get:

$$\begin{aligned}0 &= 75.0 \text{ g} \times \frac{1 \text{ cal}}{\text{g}^\circ\text{C}} \times (T - 48.7^\circ\text{C}) + 90.0 \text{ g} \times \frac{1 \text{ cal}}{\text{g}^\circ\text{C}} \times (T - 20.0^\circ\text{C}). \\-75.0 \text{ g} \times \frac{1 \text{ cal}}{\text{g}^\circ\text{C}} \times (T - 48.7^\circ\text{C}) &= 90.0 \text{ g} \times \frac{1 \text{ cal}}{\text{g}^\circ\text{C}} \times (T - 20.0^\circ\text{C}). \\-75.0T + 3650^\circ\text{C} &= 90.0T - 1800^\circ\text{C}. \\-165.0T &= -5450^\circ\text{C}. \\T &= 33.0^\circ\text{C}.\end{aligned}$$

Alternatively, a simpler method to find the equilibrium temperature of the water is to simply calculate the weighted averages of their temperatures:

$$T = \frac{75.0 \text{ g} \times 48.7^\circ\text{C} + 90.0 \text{ g} \times 20.0^\circ\text{C}}{165 \text{ g}} = 33.0^\circ\text{C}.$$

Now that we know the equilibrium temperature of the water without the reaction, we can calculate how much heat the water absorbed from the reaction. Thus, we have:

$$q_{\text{H}_2\text{O}} = mc\Delta T = 165.0 \text{ g} \times \frac{4.184 \text{ J}}{\text{g}^\circ\text{C}} \times (58.4^\circ\text{C} - 33.0^\circ\text{C}) = 17500 \text{ J}.$$

Since the heat absorbed by the water is the heat released by the reaction, we know that $q_{\text{rxn}} = -q_{\text{H}_2\text{O}} = -17500 \text{ J}$. Converting to kilojoules, we have:

$$q_{\text{rxn}} = -17500 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = -17.5 \text{ kJ}.$$

Therefore, the heat released by the reaction is $\boxed{-17.5 \text{ kJ}}$. (As the problem asks for the amount of heat released and not for ΔH , a positive answer of 17.5 kJ is also acceptable.)

(c) (No partial) The balanced chemical equation of the reaction is:



We can calculate the number of moles of HCl:

$$90.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{3 \text{ mol}}{1 \text{ L}} = 0.270 \text{ mol}.$$

We can calculate how many moles of HCl will react with 0.117 moles of Na_2S with the stoichiometric ratio to determine which reactant is limiting:

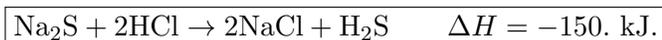
$$0.117 \text{ mol Na}_2\text{S} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Na}_2\text{S}} = 0.234 \text{ mol HCl}.$$

There are 0.270 moles of HCl, which is more than 0.234 moles, which means that Na_2S is limiting and HCl is left in excess. Thus, the amount of heat released is proportional to the amount of Na_2S in the system.

Now, we find the change in enthalpy for 1 mole of Na_2S :

$$\Delta H = 1 \text{ mol Na}_2\text{S} \times \frac{-17.5 \text{ kJ}}{0.117 \text{ mol Na}_2\text{S}} = -150. \text{ kJ}.$$

So, the two forms for the thermochemical equation would be:



Bonus Solution: (No partial)

- (a) The Second Law of Thermodynamics.
- (b) They have a lower density (greater volume) as a solid than as a liquid.
- (c) Carbon dioxide (CO_2).
- (d) From left to right: $6 + 97 + 355 + 3 + 97 + 36 + 355 + 91 + 36 = 1076$.