

CHAP 6 & 7

1. $P_{\text{apparatus}} = P_{\text{atm}} + h_1 = 785 \text{ mmHg} + 32 \text{ mmHg} = 817 \text{ mmHg}$

-SD=3

2. This is an ideal gas, so we will use this equation $PV = nRT$.

"we need P , rearrange to $P = nRT/V$

But, need n first: $n_{N_2} = \frac{29.2g}{28.02} = 1.042 \text{ mol.}$

convert $^{\circ}\text{C}$ to Kelvin.

Plug in all values! : $P = \frac{1.042 \text{ mol} \times 0.082058 \frac{\text{atm} \cdot \text{L}}{\text{K} \cdot \text{mol}} \times (273 + 71) \text{ K}}{11.4 \text{ L}}$
 $= 2.58 \text{ atm}$

3. -SD=3

The same as before, only now we do it twice. First to find initial pressure, P_i , then use ΔP , $P_e - P_i$ to find new mass. We can find mass from first finding moles.

$$P_i = \frac{n_i RT}{V}$$

$$= \frac{m_i RT}{M V}$$

$$= \frac{407 \times 0.082058 \times (23.5 + 273.15) \text{ K}}{32 \times 31}$$

$$= 9.987 \text{ atm}$$

using $P_i - P_f$, let's find n_f . We use ΔP b/c we want to find the amount of gas we want to release.

$$n_f = \frac{m_f}{M} = \frac{\Delta P V}{RT}$$

$$m = \frac{\Delta P V M}{RT} = \frac{(9.987 - 2.82)(31)(32)}{0.082058 \times (23.5 + 273.15)} = 292g$$

-SD=3

4. We know that density = $\frac{\text{mass}}{\text{Volume}}$. Because this is a gas, once again, $PV = nRT$. You notice we've got 2 unknowns, so, let's try some rearranging.

$$PV = \frac{nRT}{M} \rightarrow P = \frac{\frac{m}{M}RT}{V} \rightarrow P = \frac{dRT}{M} \rightarrow d = \frac{PM}{RT} \quad (1)$$

↳ split moles into $\frac{\text{mass}}{\text{Molar Mass}}$ that is equal to density!

Convert kPa to atm: #atm = $\frac{105 \text{ kPa}}{101.325 \text{ kPa/atm}}$
 $= 1.036 \text{ atm}$

lets use (1): $d = \frac{1.036 \text{ atm} \times 39.95 \text{ g/mol}}{0.082058 \frac{\text{atm} \cdot \text{L}}{\text{Kmol}} \times (32.1 + 273.15) \text{ K}}$

$$d = 1.65 \text{ g/L}$$

5. Whenever asked for excess reactant, find moles & get the limiting reagent!

-SD=2 $n_{\text{CO}} = \frac{PV}{RT} = \frac{1.27 \times 19}{0.082058 \times (739 + 273.15)} = 0.2905 \text{ mol}$

$$n_{\text{H}_2} = \frac{1.27 \times 40.8}{0.082058 \times (739 + 273.15)} = 0.6240 \text{ mol}$$

Assume if all CO reacted, H₂ it'd produce?

$$n_{\text{H}_2} = n_{\text{CO}} \times \frac{13}{6} = 0.2905 \times \frac{13}{6} = 0.6294 \text{ mol}$$

∴ 0.6294 > 0.6240, H₂ is limiting reagent

We need to find more than just which is limited, we need to find the amount of excess. Since we now know H₂ is limiting, let's check how much CO it can produce.

To moles: $n_{\text{CO}} = n_{\text{H}_2} \times \frac{6}{13} = 0.6240 \times \frac{6}{13} = 0.288 \text{ mol}$

Between moles: $n_{\text{excess}} = 0.2905 - 0.288 = 0.0025 \text{ mol}$

From moles: $V_{\text{excess}} = nRT/P$

$$= 0.0025 \times 0.082058 \times (739 + 273.15) / 1.27$$

$$= 0.16 \text{ L}$$

$$-SD=3$$

$$6. \quad P_{\text{atm}} = \frac{745 \text{ torr}}{760 \text{ torr/atm}} = 0.9803 \text{ atm.}$$

$$\text{To moles: } n_{\text{CO}} = \frac{PV}{RT} = \frac{0.9803 \times 28.1}{0.082058 \times (23.5 + 273.15)} = 1.132 \text{ mol}$$

$$\text{Between moles: } n_{\text{H}_2} = n_{\text{CO}} \times \frac{7}{3} = 2.6404 \text{ mol}$$

$$\text{From moles: } V_{\text{H}_2} = \frac{nRT}{P} = \frac{2.6404 \times 0.082058 \times (23.5 + 273.15)}{0.9803} = \boxed{65.6 \text{ L}}$$

$$7. \quad -SD=3.$$

$$\text{To moles: } n_{\text{LiOH}} = \frac{1110 \text{ g}}{23.95 \text{ g/mol}} = 46.35 \text{ mol}$$

$$\text{Between moles: } n_{\text{CO}_2} = n_{\text{LiOH}} \times \frac{1}{2} = 23.17 \text{ mol}$$

$$\text{From moles: } V_{\text{CO}_2} = \frac{nRT}{P} = \frac{23.17 \times 0.082058 \times (26.3 + 273.15)}{\left(\frac{763}{760}\right)}$$

$$= \boxed{567 \text{ L}}$$

$$-SD=3$$

8. This question is not difficult, more or less tedious. We'll have to find the moles of each of the products & add that up to get n_{total} .

$$\text{To moles: } n_{\text{NH}_4\text{ClO}_4} = \frac{2.41}{117.492} = 0.02051 \text{ mol}$$

$$\text{Between moles: } n_{\text{H}_2\text{O}} = n_{\text{NH}_4\text{ClO}_4} \times \frac{4}{2} = 0.04102 \text{ mol}$$

$$n_{\text{Cl}_2} = n_{\text{NH}_4\text{ClO}_4} \times \frac{1}{2} = 0.01026 \text{ mol}$$

$$n_{\text{O}_2} = n_{\text{NH}_4\text{ClO}_4} \times \frac{3}{2} = 0.03077 \text{ mol}$$

$$n_{\text{N}_2} = n_{\text{NH}_4\text{ClO}_4} \times \frac{1}{2} = 0.01026 \text{ mol}$$

$$n_{\text{gas total}} = n_{\text{H}_2\text{O}} + n_{\text{Cl}_2} + n_{\text{O}_2} + n_{\text{N}_2} = 0.08205 \text{ mol.}$$

$$\text{From moles: } P = n_{\text{gas total}} \frac{RT}{V} = 0.08205 \times 0.082058 \times (243 + 273.15) / 2.25 \text{ L}$$

$$= \boxed{1.54 \text{ atm}}$$

-SD=2

9.) When two Containers are attached, their pressures & volumes add.

$$\text{So, } P_f V_f = P_{H_2} V_{H_2} + P_{He} V_{He}$$
$$P_f = (P_{H_2} V_{H_2} + P_{He} V_{He}) / V_f$$

✓ the final Volume is equal to the sum of the volume of the two attached containers.

$$P_f = (1.5 \times 1.7 + 3.6 \times 4.8) / (4.8 + 1.7)$$
$$= 3.1 \text{ atm}$$

-SD=3
10.) Effusion and diffusion rates depend on the size of a particle. As such, molar masses should be taken into account. Use Graham's Law.

$$M_{N_2O} = 44.02 \text{ g/mol} \quad M_{NO} = 30.01 \text{ g/mol}$$

$$\frac{\text{Rate}}{\text{Ratio}} = \frac{\text{Rate}_1}{\text{Rate}_2} = \sqrt{\frac{m_2}{m_1}}$$

let us assume there are 1 mol of each molecule so we can get a rate.

$$\frac{\text{Rate}_{N_2O}}{\text{Rate}_{NO}} = \sqrt{\frac{m_{NO}}{m_{N_2O}}} = \sqrt{\frac{30.01}{44.02}} = 0.8257$$

$$\text{So, } \text{Rate}_{N_2O} = 0.8257 \text{ NO}$$

which means 1 N_2O molecule enters for every 0.8257 NO molecule.

$$\text{Plug in \& solve! } 1 \times 0.00552 \text{ mol} = 0.8257 (\text{NO})$$

$$\text{NO} = 0.00669 \text{ mol}$$

11) -SD = 3 This is a heat of Reaction type question. We'll need to use $q = mC\Delta T$

$$m_{\text{brass}} = d \times V = 8.40 \text{ g/cm}^3 \times 15.7 \text{ cm}^3 = 131.88 \text{ g}$$

$$(1) q_{\text{brass}} = 131.88 \text{ g} \times 0.385 \text{ J/g}^\circ\text{C} \times (T_2 - 174^\circ\text{C})$$

Notice the two unknowns? We'll need 2 equations. The second should be for water.

$$(2) q_{\text{water}} = 186.3 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C} \times (T_2 - 21.8^\circ\text{C}).$$

We understand that for the system to reach equilibrium, the energy lost through q_{brass} should equal the energy gained through q_{water} . Their final temperature T_2 should be equal.

$$\begin{aligned} -q_{\text{brass}} &= q_{\text{water}} \\ (131.88 \times 0.385 \times (T_2 - 174^\circ\text{C})) &= 186.3 \times 4.184 \times (T_2 - 21.8) \\ -50.77T_2 + 8834.6 &= 779.5T_2 - 16992.6 \\ 25827.3 &= 830.27T_2 \\ \boxed{T_2 = 31.1^\circ\text{C}} \end{aligned}$$

-SD = 3

12. This is a work vs pressure volume problem. So, use $w = -P_{\text{ext}} \Delta V$

$$w = -\frac{757 \text{ mmHg}}{760 \text{ mmHg/atm}} \times (31.5 - 18.5) \text{ L} \times 101.325 \frac{\text{J}}{\text{atm} \cdot \text{L}}$$

$$\boxed{= -1.31 \times 10^3 \text{ J}}$$

← This means that work was done on the surrounding

13. -SD = 3

This question is actually the same as the last one. Only now we have to solve for V_f & V_i instead of m simply being given. Use $PV = nRT$ to solve for them.

$$V_1 = 2.61 \text{ L} \quad \text{since } d_{\text{water}} = 1 \text{ g/mL}$$

$$V_2 = \frac{nRT}{P} = \frac{mRT}{MP} = \frac{2610 \times 0.082058 \times (241 + 273.15)}{18.016 \times 1} = 4923.35 \text{ L}$$

$$w = -P(V_2 - V_1)$$

$$= -1 \text{ atm} (4923.35 - 2.61) \times 101.325 \frac{\text{J}}{\text{atm} \cdot \text{L}}$$

$$= -49859.4 \text{ J}$$

$$\boxed{= -49.9 \text{ kJ}}$$

Hilroy }

$$-SD = 3$$

14. Internal Energy Change is defined by $\Delta U = q + w$.
Be careful of signs! When work is done on the system, energy is gained. When work is done on surroundings, energy lost & must be negative.

Give system gaining energy means reaction lost it

$$\Delta U = +801 + (-697) = +104 \text{ J}$$

+ive because work is done on it

$$-SD = 3$$

15. Once again, we must find ΔU , internal energy change

$$\Delta U = q + w$$

We have q already but need to solve for w

$$w = -P(V_2 - V_1)$$

$$= -0.847 \text{ atm} (11.81 \text{ L} - 4.63 \text{ L}) \times 101.325 \frac{\text{J}}{\text{atm L}} \leftarrow \text{Conversion constant}$$

$$= -616.20$$

Put everything together

$$\Delta U = +535 \text{ J} + (-616.2)$$

$$= -81.2 \text{ J}$$

$$-SD = 3$$

16. We have to find heat of combustion, which is just another way of saying enthalpy. We must find ΔH .

Using the equation: $\Delta U + \Delta nRT = \Delta H$

$$q + w + P\Delta V = \Delta H$$

remember $PV = nRT$

$$q - P\Delta V + P\Delta V = \Delta H$$

This question stated constant volume, so all $\Delta V = 0$

$$\Delta H = q_p$$

$$q = C\Delta T \leftarrow \text{Just heat capacity, not specific heat capacity, so mass}$$

$$= 4.68 \text{ kJ/g} \times (31.0 - 24.9)^\circ\text{C}$$

$$\text{We put a } \ominus \text{ there because energy was lost to surroundings. } = -28.548 \text{ kJ}$$

We've just found heat, or Enthalpy! But question wants molar Enthalpy, so, we just have to divide Enthalpy by moles! (Heat of Combustion)

$$\text{So, } \Delta H = \frac{q}{n} = \frac{-28.548 \text{ kJ}}{0.008705}$$

$$\Delta H = -3.28 \times 10^3 \text{ kJ/mol}$$

$$-SP=4$$

17. Enthalpy Change! Use: $\Delta H = \Delta U + \Delta nRT$. $\Delta n = n_f - n_i = -28$

$$\Delta U = \Delta H - (-28)(0.082058)(23+273.15)$$

$$= -1850 \text{ kJ} + 685.04 \text{ atmL} \times 101.325 \text{ J/atmL}$$

$$= -1850 \text{ kJ} + 69411.3 \text{ J}$$

$$= -1781 \text{ kJ [per mol C}_6\text{H}_{12}\text{O}]$$

look at the left side of chemical equation.
Constant Conv. atmL \rightarrow J.

18. Enthalpy of formation:

Given the Standard Enthalpy of formation for CO_2 & H_2O
We can substitute and solve for C_{10}H_8 . Diatomic molecules have H_f of 0.



$$\text{C}_{10}\text{H}_8 + 0 \frac{\text{kJ}}{\text{mol}} = 10 \times (-393.5) \frac{\text{kJ}}{\text{mol}} + 4 \times (-285.9) \frac{\text{kJ}}{\text{mol}}$$

$$\text{C}_{10}\text{H}_8 = -5078.6 \text{ kJ/mol}$$

Forming C_{10}H_8 released 5078.6 kJ for each mol of C_{10}H_8 .

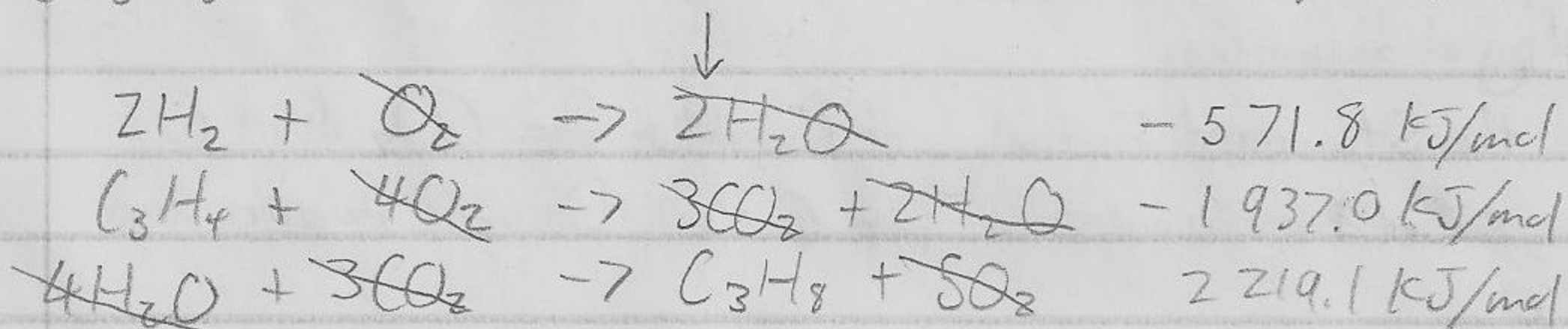
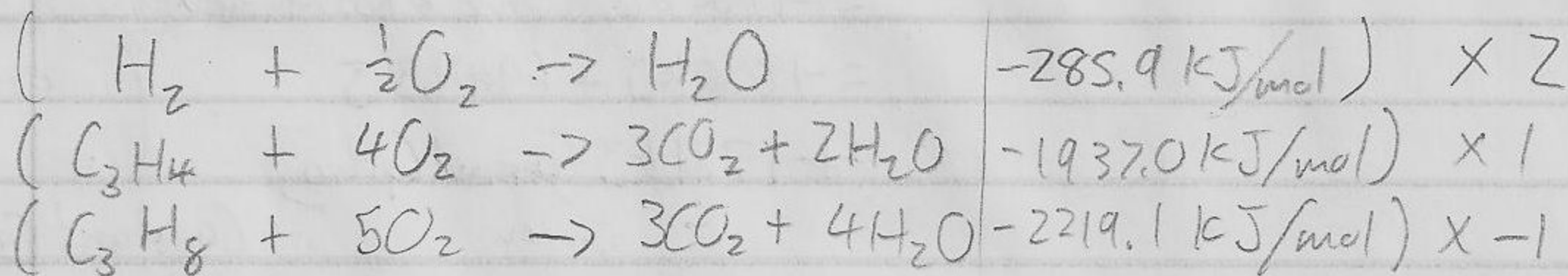
We can solve for $H_f \text{ C}_{10}\text{H}_8$ by comparing with ΔH_f

$$\Delta H_f = H_{f \text{ final}} - H_{f \text{ initial}}$$

$$-5005 \frac{\text{kJ}}{\text{mol}} = -5078.6 \frac{\text{kJ}}{\text{mol}} - H_{f \text{ initial}}$$

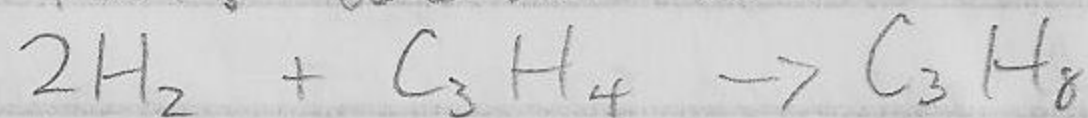
$$H_{f \text{ initial}} = \boxed{-73.6 \text{ kJ/mol}} = H_{f \text{ C}_{10}\text{H}_8}$$

19. This to me is like a puzzle piece! We have to manipulate the 3 equations above to get the target equation: $C_3H_4 + 2H_2 \rightarrow C_3H_8$



Cancel equivalent molecules.

Look at that! we now have



So, this works, then let's add up the Standard Enthalpies of Formation

$$\begin{aligned} \Delta H_{f, \text{reaction}} &= \Delta H_{f, \text{reaction 1}} + \Delta H_{f, \text{reaction 2}} + \Delta H_{f, \text{reaction 3}} \\ &= (-571.8 + (-1937) + 2219.1) \text{ kJ/mol} \\ &= \boxed{-289.7 \text{ kJ/mol}} \end{aligned}$$