

Burning of Gasoline

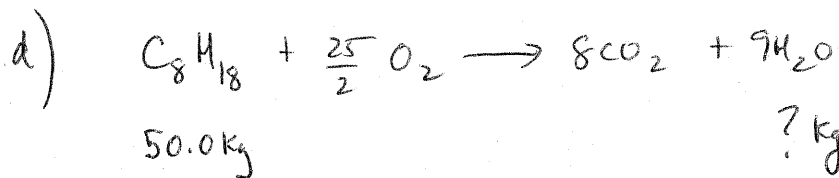
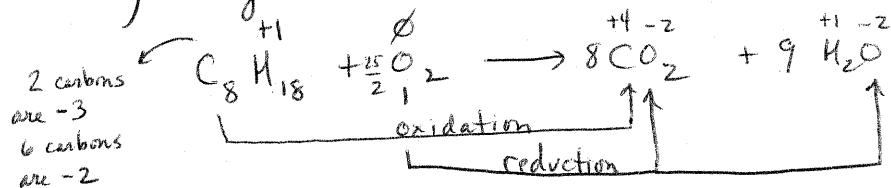


b) combustion

c) Yes, it is redox.

1) Element (O_2) becomes part of a compound (CO_2, H_2O)

2) Change in oxidation states



$$\frac{50.0 \text{ kg}}{1} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol } C_8H_{18}}{114.26 \text{ g}} \times \frac{9 \text{ mol } H_2O}{1 \text{ mol } C_8H_{18}} \times \frac{18.02 \text{ g}}{1 \text{ mol } H_2O} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 70.97 \text{ kg of } H_2O$$

(154 kg CO_2)

e) $\frac{50.0 \text{ kg}}{1} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol } C_8H_{18}}{114.26 \text{ g}} \times \frac{25 \text{ mol } O_2}{2 \text{ mol } C_8H_{18}} = 5.46998 \times 10^3 \text{ mol } O_2 = n$

$$PV = nRT \rightarrow V = \frac{nRT}{P} = \frac{(5.46998 \times 10^3 \text{ mol})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298 \text{ K})}{1.00 \text{ atm}}$$

$$= 1.3383 \times 10^5 \text{ L}$$

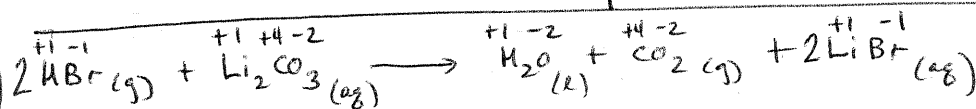
$$f) \frac{70.9 \text{ kg H}_2\text{O}}{1} \left| \frac{1000 \text{ g}}{1 \text{ kg}} \right| \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g}} = 3.93 \times 10^3 \text{ mol H}_2\text{O}$$

$$PV = nRT \rightarrow P = \frac{nRT}{V} = \frac{(3.93 \times 10^3 \text{ mol}) (0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}) 313 \text{ K}}{20.0 \text{ L}}$$

$$P = 1.01 \times 10^5 \text{ atm} \quad (8.99 \times 10^4 \text{ atm CO}_2)$$

$$\frac{1.01 \times 10^5 \text{ atm}}{1} \left| \frac{760 \text{ mmHg}}{1 \text{ atm}} \right| = 7.69 \times 10^7 \text{ mmHg}$$

$$\frac{7.69 \times 10^7 \text{ mmHg}}{1} \left| \frac{1 \text{ cmHg}}{10 \text{ mmHg}} \right| \left| \frac{1 \text{ inHg}}{2.54 \text{ cmHg}} \right| = 3.02 \times 10^6 \text{ inHg}$$



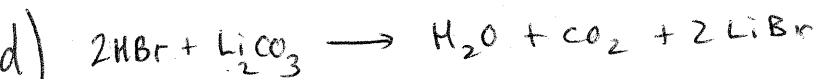
double replacement

not redox

1) DR rxns are never redox

2) oxidation states do not change, see part a

Does this make sense? Sure. 1 mol of gas occupies 22.4 L at 0°C + 1 atm. Since the T is higher and P is lower, the volume should be > (22.4 x 4)



$$P = 656 \text{ mmHg} \quad 147.8 \text{ g}$$

$$T = 42^\circ\text{C}$$

$$V = ?$$

$$n = ?$$

$$\frac{147.8 \text{ g Li}_2\text{CO}_3}{1} \left| \frac{1 \text{ mol Li}_2\text{CO}_3}{73.892 \text{ g}} \right| \frac{2 \text{ mol HBr}}{1 \text{ mol Li}_2\text{CO}_3} = 4.000 \text{ mol HBr}$$

$$PV = nRT \rightarrow V = \frac{nRT}{P} = \frac{(4.000 \text{ mol}) (0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}) (315 \text{ K})}{\frac{656 \text{ mmHg}}{760 \text{ mmHg/atm}}} = 119.86 \text{ L}$$

$$e) PV = nRT \rightarrow PV = \text{constant} \times T$$

constant

$$\frac{PV}{T} = \text{const}$$

$$\rightarrow \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$V_2 = \frac{P_1 V_1}{T_1} \cdot \frac{T_2}{P_2} = \frac{(656)(120 \text{ L})}{315 \text{ K}} \cdot \frac{293 \text{ K}}{690 \text{ mmHg}} = 106 \text{ L}$$

What is the effect on the pressure of gas sample if you simultaneously decrease its volume to 1/3 of the original volume and double its Kelvin temperature?

$PV = nRT$ What's held constant? n and R , so $PV = \text{constant} \times T$ or $\frac{PV}{T} = \text{constant}$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \rightarrow P_2 = \frac{P_1 V_1}{T_1} \cdot \frac{T_2}{V_2} \quad \text{and} \quad V_2 = \frac{V_1}{3}$$

$$T_2 = 2T_1$$

$$P_2 = \frac{P_1 V_1}{T_1} \cdot \frac{2T_1}{\frac{V_1}{3}} \rightarrow P_2 = 6P_1$$

A smaller volume and higher temperature would increase pressure.

A sample of the deadly gas, HCN, has a mass of 1.75 grams. What would be its volume at 25 °C and 750 mmHg?

1.75 g \rightarrow convert to moles $\frac{1.75 \text{ g} / 1 \text{ mol}}{27.03 \text{ g}} = 0.0647_{43} \text{ mol}$ (n)

25 °C \rightarrow convert to K $273 + 25 = 298 \text{ K}$ (T)

750 mmHg \rightarrow convert to atm $\frac{750 \text{ mmHg} / 1 \text{ atm}}{760 \text{ mmHg}} = 0.986_{54} \text{ atm}$ (P)

$PV = nRT \rightarrow V = \frac{nRT}{P}$

$$= \frac{(0.0647_{43} \text{ mol})(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})}{0.986_{54} \text{ atm}} = 1.60_{51} \text{ L}$$

A certain gas mixture contains 1.25 moles of N₂, 3.50 moles of O₂, and 5.00 moles of He. If the total pressure of the mixture is 15.0 atm, what are P_{N₂}, P_{O₂}, and P_{He}?

$P_{\text{TOTAL}} \propto n_{\text{TOTAL}}$ (Dalton's Law)

$P_T = P_{N_2} + P_{O_2} + P_{He}$ so... $n_T = n_{N_2} + n_{O_2} + n_{He}$

$$P_T = \frac{n_T RT}{V} = \frac{n_{N_2} RT}{V} + \frac{n_{O_2} RT}{V} + \frac{n_{He} RT}{V} = 1.25 + 3.50 + 5.00 = 9.75 \text{ moles}$$

$$P_{N_2} = 15.0 \text{ atm} \left(\frac{1.25}{9.75} \right) = 1.92 \text{ atm} +$$

$$P_{O_2} = 15.0 \text{ atm} \left(\frac{3.50}{9.75} \right) = 5.38 \text{ atm} +$$

$$P_{He} = 15.0 \text{ atm} \left(\frac{5.00}{9.75} \right) = 7.69 \text{ atm} +$$

$$14.99 \text{ atm} \approx 15.0 \text{ atm}$$