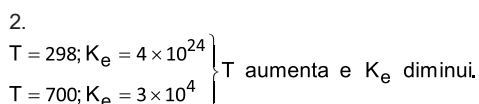
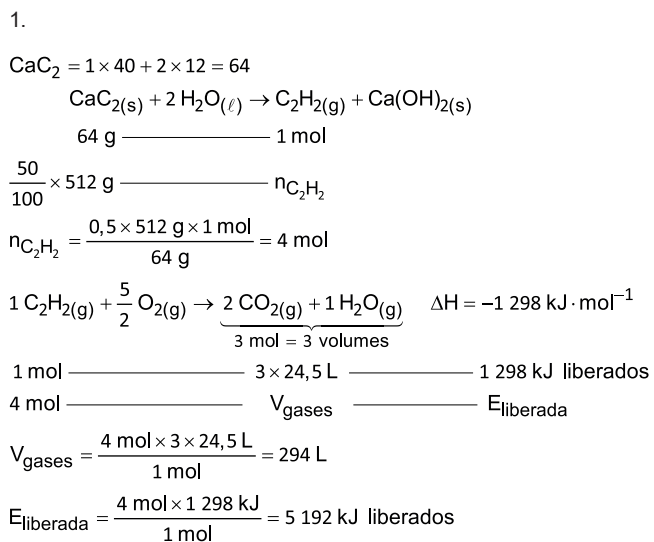
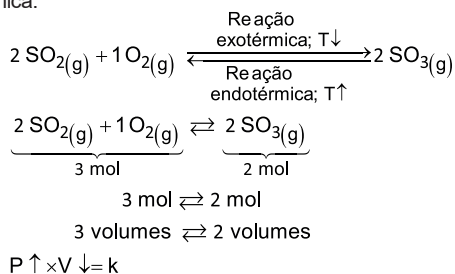




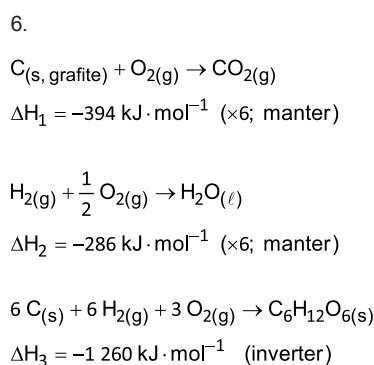
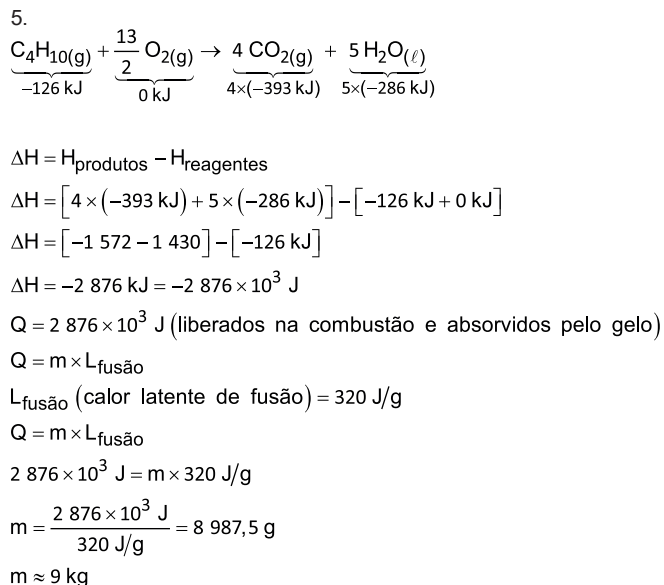
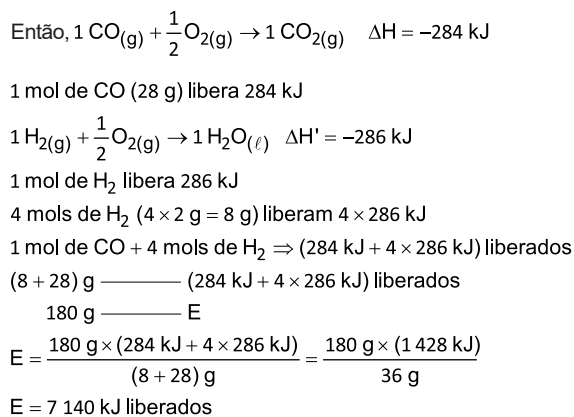
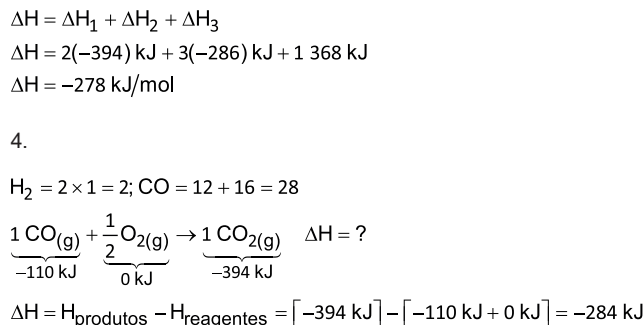
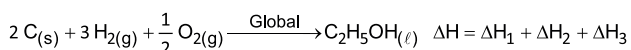
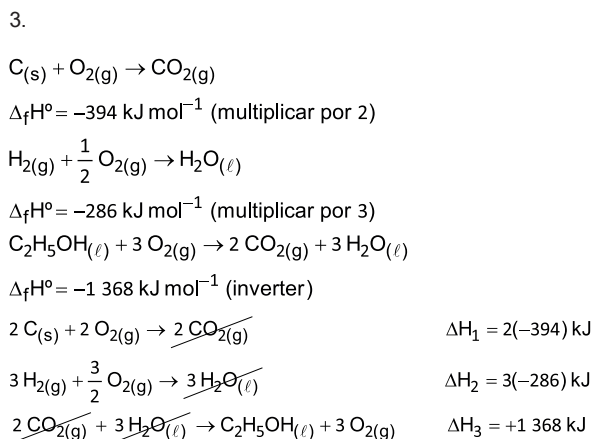
Professor: Kelton Wadson				
1	2	3	4	5
E	E	C	B	E
6	7	8	9	10
E	B	E	A	C

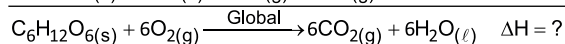
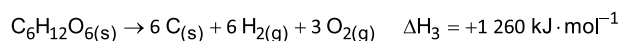


A reação direta é desfavorecida pela elevação da temperatura e favorecida pela diminuição da temperatura, ou seja, trata-se de uma reação exotérmica.



O aumento de pressão implica em deslocamento para a direita no sentido do menor volume.



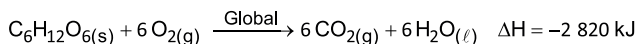


$$\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$$

$$\Delta H = [6 \times (-394) + 6 \times (-286) + 1\,260] \text{ kJ}$$

$$\Delta H = -2\,820 \text{ kJ}$$

$$\text{C}_6\text{H}_{12}\text{O}_6 = 180,0$$



$$180,0 \text{ g} \text{-----} -2\,820 \text{ kJ}$$

$$35 \text{ g} \text{-----} \text{ E}$$

$$\text{E} = \frac{35 \text{ g} \times (-2\,820 \text{ kJ})}{180,0 \text{ g}} = -548,33 \text{ kJ} \approx -548 \text{ kJ}$$

7. A entalpia de reação de hidrogenação é calculada somando-se todas as entalpias de ligação dos reagentes (que são positivas) com todas as entalpias de ligação dos produtos (que são negativas).

$$\Delta H = [4 \times (\text{C}-\text{C}) + 1 \times (\text{C}=\text{C}) + 8 \times (\text{C}-\text{H}) + 1 \times (\text{H}-\text{H})] +$$

$$[5 \times (\text{C}-\text{C}) + 10 \times (\text{C}-\text{H})]$$

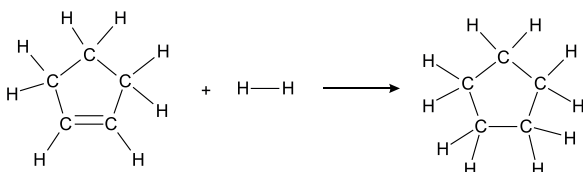
$$\Delta H = [4 \times (+335 \text{ kJ}) + 1 \times (+600 \text{ kJ}) + 8 \times (+414 \text{ kJ}) + 437 \text{ kJ}] +$$

$$[5 \times (-335 \text{ kJ}) + 10 \times (-414 \text{ kJ})]$$

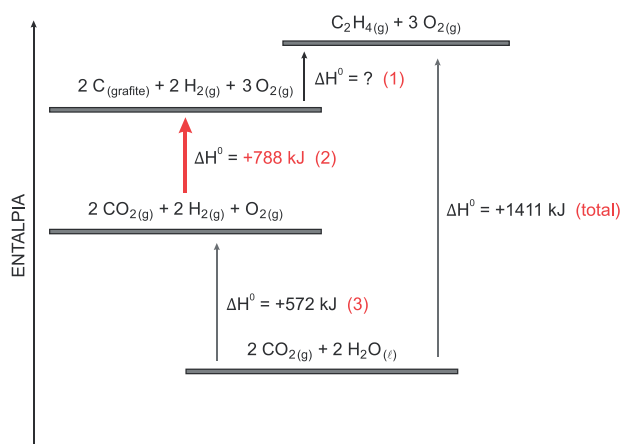
$$\Delta H = [+1\,340 \text{ kJ} + 600 \text{ kJ} + 3\,312 \text{ kJ} + 437 \text{ kJ}] +$$

$$[-1\,675 \text{ kJ} - 4\,140 \text{ kJ}]$$

$$\Delta H = +5\,689 \text{ kJ} - 5\,815 \text{ kJ} = -126 \text{ kJ/mol}$$



8.



De acordo com a Lei de Hess, temos:

$$|\Delta H_1| + |\Delta H_2| + |\Delta H_3| = \Delta H_{(\text{total})}$$

$$|\Delta H^\circ| + |-788 \text{ kJ}| + |+572 \text{ kJ}| = +1\,411 \text{ kJ}$$

$$\Delta H^\circ + 788 \text{ kJ} + 572 \text{ kJ} = +1\,411 \text{ kJ}$$

$$\Delta H^\circ + 1\,360 \text{ kJ} = +1\,411 \text{ kJ}$$

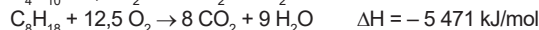
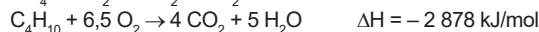
$$\Delta H^\circ = +1\,411 \text{ kJ} - 1\,360 \text{ kJ}$$

$$\Delta H^\circ = +51 \text{ kJ}$$

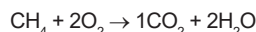
9. De acordo com a tabela,

Composto	Fórmula molecular	Massa molar (g/mol)	ΔH_{25}° (kJ/mol)
Metano	CH_4	16	-890
Butano	C_4H_{10}	58	-2 878
Octano	C_8H_{18}	114	-5 471

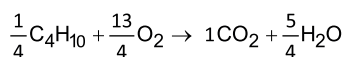
Teremos:



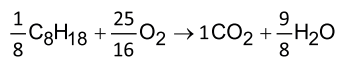
Como a comparação deve ser feita para 1 mol de CO_2 liberado por cada combustível, devemos dividir a segunda equação por dois e a terceira, por oito e, então, comparar os respectivos "novos" ΔH obtidos:



$$\Delta H = -890 \text{ kJ/mol}$$



$$\Delta H = -719,5 \text{ kJ/mol}$$



$$\Delta H = -683,875 \text{ kJ/mol}$$

Lembrando que o sinal negativo significa energia liberada, a ordem crescente de liberação será $683,875 \text{ kJ} < 719,5 \text{ kJ} < 890 \text{ kJ}$, ou seja, gasolina, GLP e gás natural.

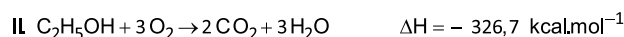
10. De acordo com a tabela, teremos as seguintes reações de combustão:



$$114 \text{ g} \text{-----} \text{ (libera) } 1\,222,5 \text{ kcal}$$

$$1 \text{ g} \text{-----} \text{ E}_\text{I}$$

$$\text{E}_\text{I} = 10,72 \text{ kcal}$$



$$46 \text{ g} \text{-----} \text{ (libera) } 326,7 \text{ kcal}$$

$$1 \text{ g} \text{-----} \text{ E}_\text{II}$$

$$\text{E}_\text{II} = 7,10 \text{ kcal}$$



$$2 \text{ g} \text{-----} \text{ (libera) } 68,3 \text{ kcal}$$

$$1 \text{ g} \text{-----} \text{ E}_\text{III}$$

$$\text{E}_\text{III} = 34,15 \text{ kcal}$$

O hidrogênio é o que apresenta menor impacto ambiental e maior vantagem energética (libera maior quantidade de energia por grama queimado).