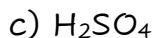
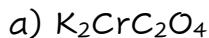


Heats of Formation, ΔH°_f

You can find the standard enthalpies of formation in the appendix of your textbook or at the following link:

Standard Enthalpies of Formation

1. Write a balanced formation equation for each the following:



2. Write a balanced formation equation for ethanol, $CH_3CH_2OH(l)$, and include the value of ΔH°_f . (look in the appendix of your book for ΔH°_f values or use the link at the top of the page).



3. Use the ΔH°_f values in the appendix of your text to determine the enthalpy change for the following reactions: (balance the equations)



$$[3 \text{ mol } CO_2 \times -393.5 \text{ kJ/mol} + 4 \text{ mol } Fe \times 0 \text{ kJ/mol}] -$$

$$[2 \text{ mol } Fe_2O_3 \times -824.2 \text{ kJ/mol} + 3 \text{ mol } C \times 0 \text{ kJ/mol}] = \mathbf{467.9 \text{ kJ}}$$



$$[1 \text{ mol } CO_2 \times -393.5 \text{ kJ/mol} + 1 \text{ mol } CaO \times -634.9 \text{ kJ/mol}] -$$

$$[1 \text{ mol } CaCO_3 \times -1207.6 \text{ kJ/mol}] = \mathbf{+179.2 \text{ kJ}}$$

4. Use bond dissociation energies to calculate ΔH°_{rxn} for the following reaction.



Reactants; 2 H–O and 3 S=O Product: 2 S–O, 2 S=O, 2 H–O

$$[2 \text{ mols } H-O \times 463 \text{ kJ/mol} + 3 \text{ mol } S=O \times 523 \text{ kJ/mol}] -$$

$$2 \text{ mol } S-O \times 364 \text{ kJ/mol} + 2 \text{ mol } S=O \times 523 \text{ kJ/mol} + 2 \text{ mol } H-O \times$$

$$463 \text{ kJ/mol} = \mathbf{-205 \text{ kJ}}$$