

Enthalpy

In any chemical reaction, the side having **minimum enthalpy** (minimum energy) is said to be favoured because molecules tend to go to or remain at a state of minimum energy. Thus, there is a tendency for **exothermic reactions** to be **spontaneous**. Note that a spontaneous change is a change which occurs by itself, without outside assistance.

For each of the following reactions, decide whether the **reactants** or the **products** have the lesser enthalpy (energy) and would therefore be favoured:

- $2\text{N}_2\text{O}(\text{g}) \rightarrow 2\text{N}_2(\text{g}) + \text{O}_2(\text{g}) + 164 \text{ kJ}$ products
- $2\text{CO}(\text{g}) \rightarrow 2\text{C}(\text{s}) + \text{O}_2(\text{g}) \quad \Delta H = 221 \text{ kJ}$ reactants
- $2\text{HCl}(\text{g}) + 184 \text{ kJ} \rightarrow \text{H}_2(\text{g}) + \text{Cl}_2(\text{g})$ reactants
- $2\text{NO}(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{NOCl}(\text{g}) \quad \Delta H = -76 \text{ kJ}$ products
- $\text{C}_2\text{H}_2\text{Cl}_4(\text{l}) \rightarrow \text{C}_2\text{H}_2(\text{g}) + 2\text{Cl}_2(\text{g}) \quad \Delta H = 386 \text{ kJ}$ reactants
- $\text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) + 394 \text{ kJ}$ products

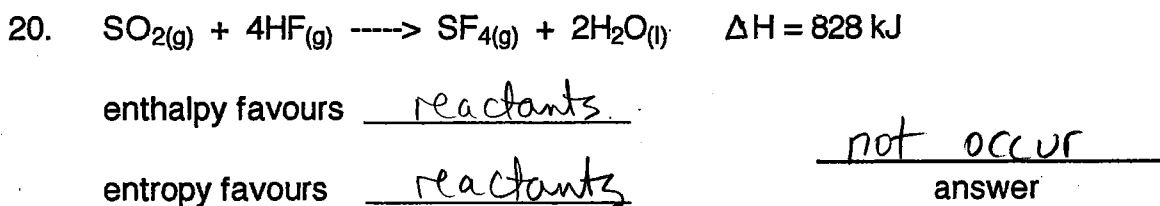
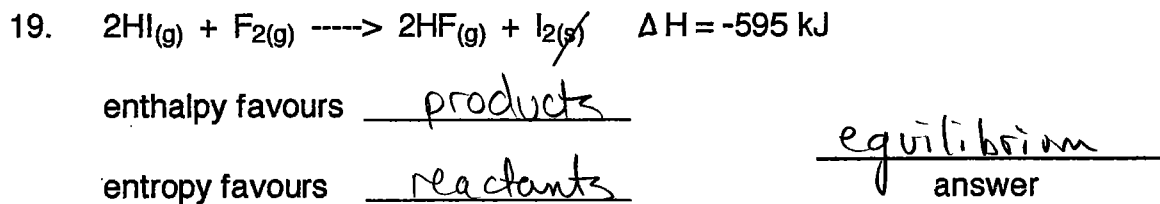
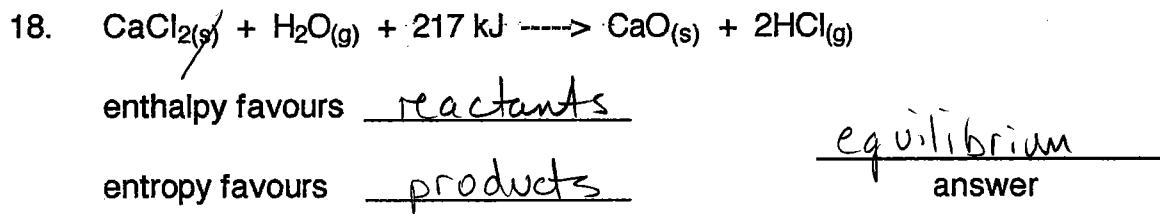
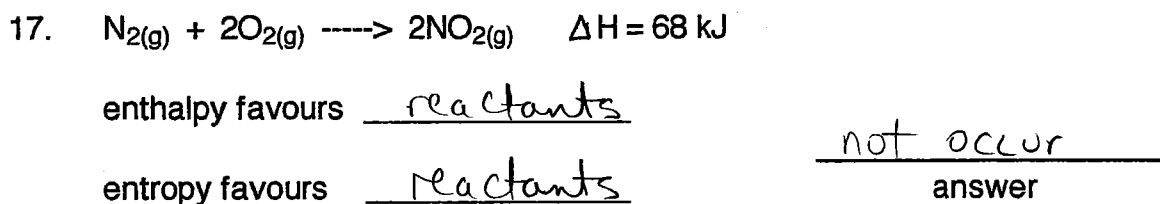
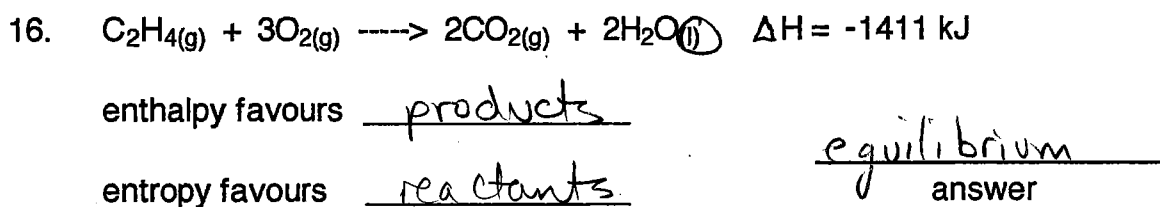
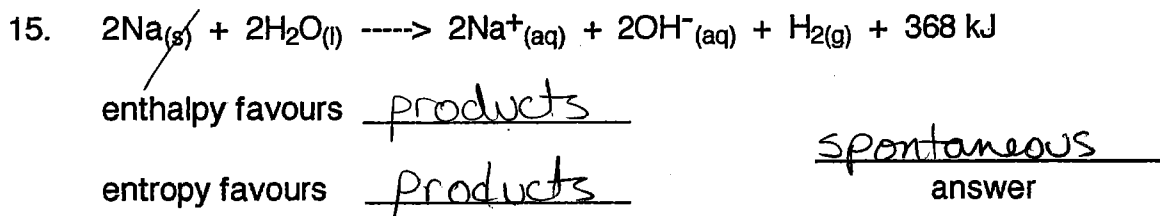
Entropy

In all chemical and physical changes there is a tendency for the system to attain a state of **maximum entropy** or randomness. The randomness of phases from greatest to least is: **gases >> solutions > liquids >> solids**. In general, the side of the equation favoured is the side having the most particles of the most random phase.

For each of the following reactions, decide whether the **reactants** or the **products** have the greater entropy and would therefore be favoured. There may also be **no significant change**.

- $\text{Na}_2\text{CO}_3(\text{s}) \rightarrow 2\text{Na}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$ products
- $\text{H}_2\text{O}(\text{s}) \rightarrow \text{H}_2\text{O}(\text{l})$ products
- $\text{CaO}(\text{s}) + \text{SO}_2(\text{g}) \rightarrow \text{CaSO}_3(\text{s})$ reactants
- $2\text{HI}(\text{g}) \rightarrow \text{H}_2(\text{g}) + \text{I}_2(\text{g})$ no significant change
- $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$ reactants
- $2\text{Al}(\text{s}) + 3\text{Br}_2(\text{l}) \rightarrow 2\text{AlBr}_3(\text{s})$ reactants
- $2\text{N}_2\text{O}_5(\text{g}) \rightarrow 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$ products
- $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$ reactants

For each of the following reactions, use enthalpy and entropy considerations to decide whether a reaction in the direction shown will be **spontaneous**, reach a state of **equilibrium**, or **not occur**. Assume that the system is closed.



This is different: Predict whether the equilibrium $4\text{HCl}_{(g)} + \text{O}_{2(g)} \xrightarrow{\Delta S} 2\text{H}_2\text{O}_{(g)} + 2\text{Cl}_{2(g)}$ is exothermic or endothermic and explain your reasoning. $\Delta H = \rightarrow$

exothermic! To be at equilibrium the products must be favoured to be opposite the ΔS (reverse) reaction.