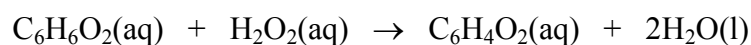
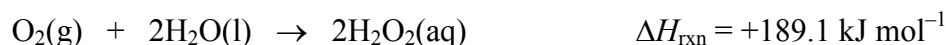
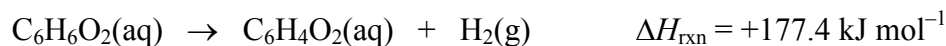


Marks
8

- The conversion of hydroquinone ($C_6H_6O_2(aq)$) to quinone ($C_6H_4O_2(aq)$) is involved in many important biochemical reactions. The bombardier beetle, for example, uses the explosive reaction between hydroquinone and hydrogen peroxide (as described by the equation below) as a defence mechanism.



From the following reaction data, calculate ΔH_{rxn} for the reaction between 1.00 mol of hydroquinone and 1.00 mol of hydrogen peroxide.

 $\Delta H_{rxn} =$

Use the answer you obtained above to calculate the heat liberated (in joules) in the oxidation of 3.86×10^{-4} mol of hydroquinone to quinone.

Answer:

Calculate the temperature rise of 0.250 g of water for this quantity of heat.
(The heat capacity of water, $C_p = 4.184 \text{ J K}^{-1} \text{ g}^{-1}$)

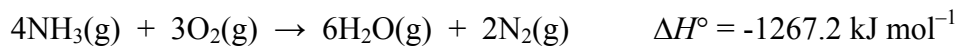
Answer:

Marks
6

- The final step in the industrial production of urea, $\text{CO}(\text{NH}_2)_2$, is:



Using the following data, calculate the standard enthalpy of formation ΔH°_f of solid urea.

 $\Delta H^\circ_f =$

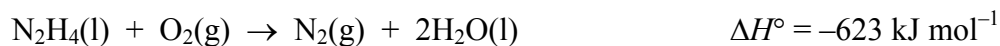
The formation of urea in this process is only spontaneous above 821°C . What is the value of the entropy change ΔS° (in $\text{J K}^{-1} \text{mol}^{-1}$) for the reaction?

 $\Delta S^\circ =$

Rationalise the sign of ΔS° in terms of the physical states of the reactants and products.

Marks
3

- The combustion of hydrazine, N_2H_4 , with oxygen is described by the following equation:



Given that ΔH°_f of $\text{H}_2\text{O}(\text{l})$ is -286 kJ mol^{-1} , find the standard enthalpy of formation of $\text{N}_2\text{H}_4(\text{l})$.

 $\Delta H^\circ_f =$

The combustion of 1.00 mol of $\text{N}_2\text{H}_4(\text{l})$ can also be accomplished using $\text{N}_2\text{O}_4(\text{l})$ as the oxidant, whereupon 629 kJ of energy is released at standard temperature and pressure. What is the standard enthalpy of formation of $\text{N}_2\text{O}_4(\text{l})$?

 $\Delta H^\circ_f =$

Marks
6

- High-purity benzoic acid, $\text{C}_6\text{H}_5\text{COOH}$, ($\Delta H^\circ_{\text{comb}} = -3227 \text{ kJ mol}^{-1}$) is used to calibrate a bomb calorimeter that has a 1.000 L capacity. A 1.000 g sample of $\text{C}_6\text{H}_5\text{COOH}$ is placed in the bomb calorimeter, along with 750 mL of pure $\text{H}_2\text{O}(\text{l})$, and the remaining 250 mL cavity is filled with pure $\text{O}_2(\text{g})$ at 10.00 atm. The $\text{C}_6\text{H}_5\text{COOH}$ is ignited and completely burned, causing the temperature of the water and the bomb calorimeter to rise from 27.20 °C to 33.16 °C. Write the chemical equation corresponding to the standard enthalpy of combustion ($\Delta H^\circ_{\text{comb}}$) of $\text{C}_6\text{H}_5\text{COOH}$.

Given that $\text{H}_2\text{O}(\text{l})$ has a heat capacity of $4.184 \text{ J K}^{-1} \text{ g}^{-1}$ and a density of 0.997 g mL^{-1} , calculate the heat capacity of the bomb calorimeter itself (in units of J K^{-1}). Ignore the heat capacity of the gases and of $\text{C}_6\text{H}_5\text{COOH}$.

Answer:

If 30.0% of the CO_2 produced dissolves in the water, calculate the final total pressure (in atm) inside the 250 mL cavity of the bomb calorimeter. Assume oxygen is insoluble in water and ignore the vapour pressure of water.

Answer:

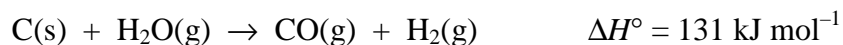
- The specific heat capacity of water is $4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and the specific heat capacity of copper is $0.39 \text{ J g}^{-1} \text{ K}^{-1}$. If the same amount of energy were applied to a 1.0 mol sample of each substance, both initially at $25 \text{ }^\circ\text{C}$, which substance would get hotter? Show all working.

Marks
2

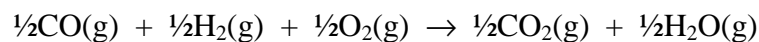
Answer:

Marks
3

- “Water gas” is a mixture of combustible gases produced from steam and coal according to the following reaction:



The equation for the complete combustion of 1 mol of water gas (*i.e.* 0.5 mol CO(g) and 0.5 mol H₂(g)) can be written as:



Calculate the standard enthalpy of combustion of water gas, given the following thermochemical data.

$$\Delta H^\circ_{\text{vap}}(\text{H}_2\text{O}) = 44 \text{ kJ mol}^{-1}$$

$$\Delta H^\circ_{\text{f}}(\text{H}_2\text{O(l)}) = -286 \text{ kJ mol}^{-1}$$

$$\Delta H^\circ_{\text{f}}(\text{CO}_2\text{(g)}) = -393 \text{ kJ mol}^{-1}$$

Answer:

THIS QUESTION CONTINUES ON THE NEXT PAGE.

Marks
4

- A solution of 2.00 M NaOH (50.0 mL) at 44.9 °C is added to a constant pressure (“coffee cup”) calorimeter containing 250.0 mL of 0.70 M HNO₃ at 21.5 °C. The final temperature of the solution is 29.9 °C. Calculate the enthalpy of neutralisation of OH⁻(aq) and H⁺(aq) in kJ mol⁻¹. Assume the density of these solutions is 1.000 g mL⁻¹ and the specific heat capacity of the solutions is 4.184 J K⁻¹ g⁻¹.

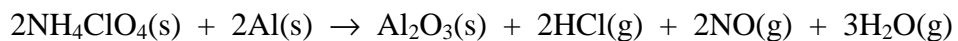
Answer:

Calculate the pH in the combined solution in the calorimeter at 21.5 °C.

Answer:

Marks
3

- Ammonium perchlorate mixed with powdered aluminium powers the space shuttle booster rockets:



Given the following thermochemical data, how much heat would be released per gram of Al(s)?

$$\Delta H_f^\circ (\text{H}_2\text{O}(\text{l})) = -285.1 \text{ kJ mol}^{-1}$$

$$\Delta H_f^\circ (\text{NO}(\text{g})) = 90.4 \text{ kJ mol}^{-1}$$

$$\Delta H_f^\circ (\text{HCl}(\text{g})) = -92.3 \text{ kJ mol}^{-1}$$

$$\Delta H_f^\circ (\text{Al}_2\text{O}_3(\text{s})) = -1669.8 \text{ kJ mol}^{-1}$$

$$\Delta H_f^\circ (\text{NH}_4\text{ClO}_4(\text{s})) = -290.6 \text{ kJ mol}^{-1}$$

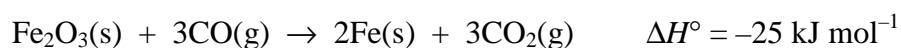
$$\Delta H_{\text{vap}}^\circ (\text{H}_2\text{O}) = 44.1 \text{ kJ mol}^{-1}$$

Answer:

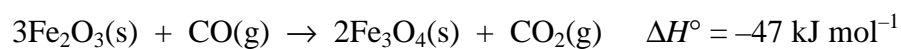
THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.

Marks
5

- Carbon monoxide is commonly used in the reduction of iron ore to iron metal. Iron ore is mostly haematite, Fe_2O_3 , in which case the complete reduction reaction is:



Incomplete reduction, however, results in the formation of magnetite, Fe_3O_4 :



Use these heats of reaction to calculate the enthalpy change when one mole of magnetite is reduced to iron metal using carbon monoxide.

Answer:

Another iron oxide that can be formed as an intermediate during reduction is FeO. Use the following table of thermochemical data to show whether the formation of FeO from Fe_3O_4 is spontaneous or not at 25 °C.

	$\Delta_f H^\circ$ (kJ mol ⁻¹)	S° (J K ⁻¹ mol ⁻¹)
FeO	-272	61
Fe ₃ O ₄	-1118	146
CO	-111	198
CO ₂	-394	214

Marks
4

- A 150.0 g block of iron metal is cooled by placing it in an insulated container with a 50.0 g block of ice at 0.0 °C. The ice melts, and when the system comes to equilibrium the temperature of the water is 78.0 °C. What was the original temperature (in °C) of the iron?

Data: The specific heat capacity of liquid water is $4.184 \text{ J K}^{-1} \text{ g}^{-1}$.

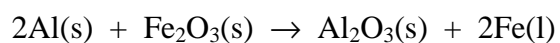
The specific heat capacity of solid iron is $0.450 \text{ J K}^{-1} \text{ g}^{-1}$.

The molar enthalpy of fusion of ice (water) is $6.007 \text{ kJ mol}^{-1}$.

Answer:

Marks
6

- The thermite reaction is written below. Show that the heat released in this reaction is sufficient for the iron to be produced as molten metal.



Assume that the values in the table are independent of temperature.

Substance	Enthalpy of formation, $\Delta_f H^\circ$ kJ mol ⁻¹	Molar heat capacity, C_p J K ⁻¹ mol ⁻¹	Melting point °C	Enthalpy of fusion kJ mol ⁻¹
Al	0	24	660	11
Al ₂ O ₃	-1676	79	2054	109
Fe	0	25	1535	14
Fe ₂ O ₃	-824	104	1565	138