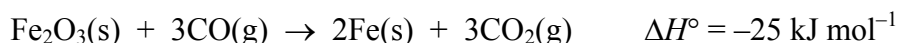
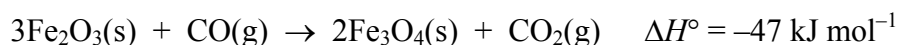


Marks
5

- 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:

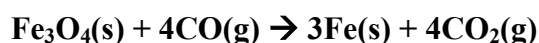


Incomplete reduction, however, results in the formation of magnetite, Fe_3O_4 :

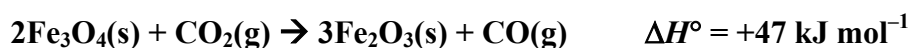


Use these heats of reaction to calculate the enthalpy change when one mole of magnetite is reduced to iron metal using carbon monoxide.

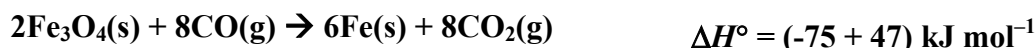
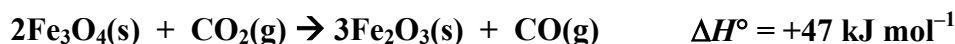
The required reaction is:



The second reaction in the question is reversed so that it leads to loss of $\text{Fe}_3\text{O}_4(\text{s})$:



This reaction is then added to $3 \times$ the first reaction:



The chemical reaction is exactly twice that required, so for one mole of $\text{Fe}_3\text{O}_4(\text{s})$, the $\Delta H^\circ = (-75 + 47) / 2 \text{ kJ mol}^{-1} = -14 \text{ kJ mol}^{-1}$.

Alternatively, using the data in the next part of the question,

$$\Delta_{\text{rxn}}H^\circ = \sum m\Delta_fH^\circ(\text{products}) - \sum n\Delta_fH^\circ(\text{reactants}),$$

$$\Delta_{\text{rxn}}H^\circ = [4\Delta_fH^\circ(\text{CO}_2(\text{g}))] - [\Delta_fH^\circ(\text{Fe}_3\text{O}_4(\text{s})) + 4\Delta_fH^\circ(\text{CO}(\text{g}))]$$

$\Delta_fH^\circ(\text{Fe}(\text{s})) = 0$ as it is an element in its standard state.

Hence using the data in the table below:

$$\Delta_{\text{rxn}}H^\circ = ([4 \times -394] - [-1118 + 4 \times -111]) \text{ kJ mol}^{-1} = -14 \text{ kJ mol}^{-1}$$

Answer: $\Delta_{\text{rxn}}H^\circ = -14 \text{ kJ mol}^{-1}$

ANSWER CONTINUES ON THE NEXT PAGE

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 Fe₃O₄ 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

For the reaction,



$$\begin{aligned}\Delta_{\text{rxn}}H^\circ &= \sum m\Delta_f H^\circ(\text{products}) - \sum n\Delta_f H^\circ(\text{reactants}) \\ &= ([3 \times -272 - 394] - [-1118 - 111]) \text{ kJ mol}^{-1} = +19 \text{ kJ mol}^{-1}\end{aligned}$$

$$\begin{aligned}\Delta_{\text{rxn}}S^\circ &= \sum mS^\circ(\text{products}) - \sum nS^\circ(\text{reactants}) \\ &= ([3 \times 61 + 214] - [146 + 198]) \text{ J K}^{-1} \text{ mol}^{-1} = +53 \text{ J K}^{-1} \text{ mol}^{-1}\end{aligned}$$

Thus,

$$\begin{aligned}\Delta_{\text{rxn}}G^\circ &= \Delta_{\text{rxn}}H^\circ - T\Delta_{\text{rxn}}S^\circ \\ &= (+19 \times 10^3 \text{ J mol}^{-1}) - (298 \text{ K})(53 \text{ J K}^{-1} \text{ mol}^{-1}) \\ &= +3200 \text{ J mol}^{-1} = +3.2 \text{ kJ mol}^{-1}\end{aligned}$$

As $\Delta_{\text{rxn}}G^\circ > 0$, the reaction is not spontaneous.