**Marks
5**
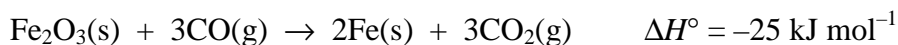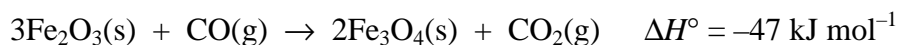
- Carbon monoxide is commonly used in the reduction of iron ore to iron metal. Iron ore is mostly haematite, $Fe_2O_3$, in which case the complete reduction reaction is:

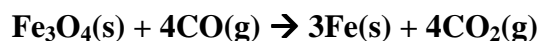$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g) \qquad \Delta H° = -25 \text{ kJ mol}^{-1}$$

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

$$3Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g) \qquad \Delta H° = -47 \text{ kJ mol}^{-1}$$
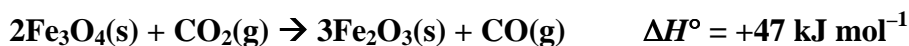
Use these heats of reaction to calculate the enthalpy change when one mole of magnetite is reduced to iron metal using carbon monoxide.
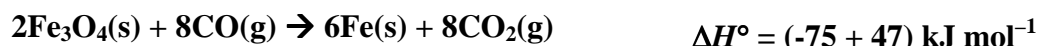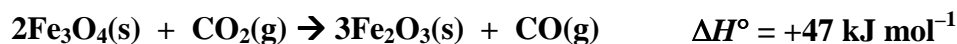
**The required reaction is:**

**$Fe_3O_4(s) + 4CO(g) \rightarrow 3Fe(s) + 4CO_2(g)$**

**The second reaction in the question is reversed so that it leads to loss of $Fe_3O_4(s)$:**

**$2Fe_3O_4(s) + CO_2(g) \rightarrow 3Fe_2O_3(s) + CO(g) \qquad \Delta H° = +47 \text{ kJ mol}^{-1}$**

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

| | |
|---|---|
| **$3Fe_2O_3(s) + 9CO(g) \rightarrow 6Fe(s) + 9CO_2(g)$** | **$\Delta H° = 3 \times -25 \text{ kJ mol}^{-1}$** |
| **$2Fe_3O_4(s) + CO_2(g) \rightarrow 3Fe_2O_3(s) + CO(g)$** | **$\Delta H° = +47 \text{ kJ mol}^{-1}$** |
| **$2Fe_3O_4(s) + 8CO(g) \rightarrow 6Fe(s) + 8CO_2(g)$** | **$\Delta H° = (-75 + 47) \text{ kJ mol}^{-1}$** |

**The chemical reaction is exactly twice that required, so for one mole of $Fe_3O_4(s)$, the $\Delta H° = (-75 + 47) / 2 \text{ kJ mol}^{-1} = -14 \text{ kJ mol}^{-1}$.**

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

**$\Delta_{rxn}H° = \Sigma m\Delta_f H°(\text{products}) - \Sigma n\Delta_f H°(\text{reactants})$,**

**$\Delta_{rxn}H° = [4\Delta_f H°(CO_2(g)] - [\Delta_f H°(Fe_3O_4(s) + 4\Delta_f H°(CO(g)]$**

**$\Delta_f H°(Fe(s)) = 0$ as it is an element in its standard state.**

**Hence using the data in the table below:**

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

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

**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 $Fe_3O_4$ is spontaneous or not at 25 °C.

|  | $\Delta_fH°$ (kJ mol$^{-1}$) | $S°$ (J K$^{-1}$ mol$^{-1}$) |
|---|---|---|
| FeO | –272 | 61 |
| $Fe_3O_4$ | –1118 | 146 |
| CO | –111 | 198 |
| $CO_2$ | –394 | 214 |

**For the reaction,**

$$Fe_3O_4(s) + CO(g) \rightarrow 3FeO(s) + CO_2(g)$$

$$\Delta_{rxn}H° = \Sigma m\Delta_fH°(products) - \Sigma n\Delta_fH°(reactants)$$
$$= ([3 \times -272 -394] – [-1118 – 111])\ kJ\ mol^{-1} = +19\ kJ\ mol^{-1}$$

$$\Delta_{rxn}S° = \Sigma mS°(products) - \Sigma nS°(reactants)$$
$$= ([3 \times 61 + 214] – [146 + 146])\ J\ K^{-1}\ mol^{-1} = +53\ J\ K^{-1}\ mol^{-1}$$

**Thus,**

$$\Delta_{rxn}G° = \Delta_{rxn}H° -T\Delta_{rxn}S°$$
$$= (+19 \times 10^3\ J\ mol^{-1}) – (298\ K)(53\ J\ K^{-1}\ mol^{-1})$$
$$= +3200\ J\ mol^{-1} = +3.2\ kJ\ mol^{-1}$$

**As $\Delta_{rxn}G° > 0$, the reaction is not spontaneous.**