- A saturated solution of lithium carbonate in pure water at 20 °C contains 1.33 g of solute per 100.0 mL of solution. Calculate the aqueous solubility product of lithium carbonate at this temperature. **4**

The molar mass of Li_2CO_3 is (2 × 6.941 (Li) + 12.01 (C) + 3 × 16.00 (O)) g mol^{-1} = 73.892 g mol^{-1}. A mass of 1.33 g therefore corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{1.33 \text{ g}}{73.892 \text{ g mol}^{-1}} = 0.0180 \text{ mol}$$

The reaction table for the dissolution of Li_2CO_3 is:

	Li_2CO_3	$\rightleftharpoons$	$2Li^+(aq)$	$CO_3^{2-}(aq)$
Initial	**0.0180**		**0**	**0**
Change	**-x**		**+2x**	**+x**
Equilibrium	**-**		**0.0360**	**0.0180**

These number of moles of $Li^+(aq)$ and $CO_3^{2-}(aq)$ in 100.0 mL. In a litre, the concentrations are therefore $[Li^+(aq)]$ = 0.360 M and $[CO_3^{2-}(aq)]$ = 0.180 M. The solubility product is therefore:

$$K_{sp} = [Li^+(aq)]^2[CO_3^{2-}(aq)] = (0.360)^2(0.180) = 0.0233$$

K_{sp} = **0.0233**

When the temperature of the same solution is raised to 40 °C, the solubility is reduced to 1.17 g per 100.0 mL of solution. What conclusions can be drawn about the sign of the standard enthalpy of dissolution of lithium carbonate?

Increasing the temperature leads to less dissolution: the equilibrium has shifted towards reactants (to the left). According to Le Chatelier's principle, this is consistent with an exothermic reaction: $\Delta H < 0$.