The equation above is to calculate the entropy change of surroundings. This involves dividing the enthalpy change (in J mol^{-1}) by the temperature and turning this value negative.

The unit for entropy change is *always* J K^{-1} mol^{-1}. Never kJ K^{-1} mol^{-1} which explains why the units for the enthalpy change (when calculating the entropy change of surroundings) must always be J mol^{-1}

Example:

Calculate the total entropy change for the reaction of ammonia and hydrogen chloride under standard conditions.

$$NH_{3(g)} + HCl_{(g)} \rightarrow NH_4Cl_{(s)}$$

$\Delta H = -315$ kJ mol^{-1} (at 298K)
$S[NH_{3(g)}] = 192.3$ J K^{-1} mol^{-1}
$S[HCl_{(g)}] = 186.8$ J K^{-1} mol^{-1}
$S[NH_4Cl_{(s)}] = 94.6$ J K^{-1} mol^{-1}

First off, find the entropy change of the system;

We know that from the previous equation to calculate the entropy change of system, we need to know the total entropy of the reactants and the products. There are two reactants so we add the entropies together and one product so we don't do anything to that.

$$S_{reactants} = 192.3 + 186.8 = \mathbf{379.1}$$
$$S_{products} = \mathbf{94.6}$$

Now we work out the difference between the values:

$$S_{system} = 94.6 - 376.1 = -284.5 \text{ J K}^{-1} \text{ mol}^{-1}$$

Now, we have to consider calculating the entropy change of the surroundings

$$\Delta S_{surroundings} = -\frac{\Delta H}{T}$$

The value for ΔH is -315 kJ mol^{-1} (at 298K, this is important information). This needs to be converted so we multiply it by 1000 to change the units from kJ mol^{-1} to J mol^{-1}.

$-315 \times 1000 = -315 \times 10^3$ J mol^{-1}

Notice the exam question has mentioned "under standard conditions" which implies that the temperature is 298K. So the value of T is 298