540 calories of heat is required to vaporize 1 gm of water at 100℃. Determine the entropy change involved in vaporizing 5 gm of water?

**Solution.**

$$L = 540\frac{cal}{g}, t = 100℃, m = 5g = 0.005kg;$$

$$\Delta S - ?$$

$L = 540\frac{cal}{g}$ - the latent heat of vaporize of water.

Converting the latent heat to joules per grams:

$$1\ cal = 4.1868J;$$

$$L = 540\frac{cal}{g}(4.1868J) = 2260.872\frac{J}{g}.$$

Converting the latent heat to joules per kilograms:

$$L = 2260.872\frac{J}{g}\left(\frac{1000g}{1kg}\right) = 2260872\frac{J}{kg}.$$

$$L = 2260872\frac{J}{kg}.$$

The entropy change is:

$$\Delta S = \frac{\Delta Q}{T}.$$

$\Delta Q$ – the heat that is required to vaporize *5 g* of water.

$$\Delta Q = mL.$$

Thermodynamic temperature:

$$T = (t + 273℃)K;$$

$$t = 100℃;$$

$$T = (100℃ + 273℃)K;$$

$$T = 373K.$$

$$\Delta S = \frac{mL}{T}.$$

$$\Delta S = \frac{0.005kg \cdot 2260872\frac{J}{kg}}{373K} = 30.3\frac{J}{K}.$$

**Answer:** The entropy change is $\Delta S = 30.3\frac{J}{K}$.