UV-VISIBLE SPECTROSCOPY – THE BEER-LAMBERT LAW

1. a) A is the absorbance of the sample.

b) I_0 and I are the intensities of light passing through the liquid under investigation. In the case of I_0, the liquid is the pure solvent. In the case of I, the liquid is a solution of the substance you are studying in the same solvent.

Light of a particular wavelength is passed through both liquids separately in identical cells, and the intensity of light in each case is measured. That is then repeated for all the other wavelengths that the device covers in the UV and visible spectrum.

c) I_0/I would be 100/50 = 2. $\log_{10}2 = 0.30$

d) l is the length of the solution that the light is passing through, and c is its concentration. Light is absorbed as it interacts with the molecules in the solution – the more molecules there are, the more light will be absorbed. Increasing the length of the solution the light passes through or the concentration of the solution will both increase the number of molecules that the light interacts with.

e) A has no units – it is the logarithm of a ratio.

That means that the expression εlc can't have any units either.

But l has the units cm, and c has the units mol dm^{-3}.

That means that the units of ε must exactly cancel those out: cm^{-1} mol^{-1} dm^3. (If you have written these in a different order, it doesn't matter.)

f) It shows the two wavelengths absorbed by electron jumps involving the oxygen lone pair and the pi electrons in the C=O double bond in ethanal. The molar absorptivity values show that the absorption at 180 nm is very much stronger than the one at 290 nm.

g) If you plotted the molar absorptivity values from the previous question as they are, the scale of the graph would be such that the absorption at 290 nm would be virtually undetectable if your scale had to go up to 10000. By plotting log values, the numbers turn out to be 1.18 and 4, which fit easily on the same axis.