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### 3.1.6 Chemical equilibria, Le Chatelier's principle and Kc ...

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The first step is to replace the molecules in the square brackets with the units of concentration. Using our previous example:  $K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(\text{mol dm}^{-3})^2}{(\text{mol dm}^{-3})(\text{mol dm}^{-3})}$  You can see that the numerator,  $(\text{mol dm}^{-3})^2$ , is the same as the denominator,  $(\text{mol dm}^{-3})(\text{mol dm}^{-3})$  and so all the units cancel.

### How do you calculate the equilibrium constant, Kc, of a ...

Calculate Kc at this temperature.  $\text{CH}_3\text{COOH}(\text{l}) + \text{CH}_3\text{CH}_2\text{OH}(\text{l}) \rightleftharpoons \text{CH}_3\text{COOCH}_2\text{CH}_3(\text{l}) + \text{H}_2\text{O}(\text{l})$  .1.00 a-. 00. ..v.s.s.s o. XY<S. +0. (4) (A2 BBoF 29) 1.90 mol of hydrogen and 1.90 mol of iodine were allowed to reach equilibrium at 710 K. The equilibrium mixture contained 3.00 mol of HI.

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3of 0.100 mol dm<sup>-3</sup>sodium butanoate. c) 1.00 g of potassium ethanoate is dissolved in 50 cm<sup>3</sup>of 0.200 mol dm<sup>-3</sup>ethanoic acid (K. a= 1.74 x 10. <sup>-5</sup>mol dm<sup>-3</sup>). 2) Calculate the pH of the following buffer solutions made by mixing an excess of weak acids with strong bases. a) 25 cm<sup>3</sup>of 0.5 mol dm<sup>-3</sup>methanoic acid (K.

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