Assignment: sample_equilibrium

(a) ICE Table for $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$:

	N_2	H_2	NH_3
Initial (M)	0.500	0.800	0
Change (M)	$-x$	$-3x$	$+2x$
Equilibrium (M)	$0.500 - x$	$0.800 - 3x$	$2x$

(b) The equilibrium constant expression is:

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Plugging in the equilibrium values:

$$0.500 = \frac{(2x)^2}{(0.500 - x)(0.800 - 3x)^3}$$

Rearranging the equation:

$$0.500 = \frac{4x^2}{(0.500 - x)(0.800 - 3x)^3}$$

$$0.500(0.500 - x)(0.800 - 3x)^3 = 4x^2$$

Solving for x by successive approximation or numerical root-finding. For $x = 0.0846$:

$$Q = \frac{(2 \cdot 0.0846)^2}{(0.500 - 0.0846)(0.800 - 3 \cdot 0.0846)^3} = \frac{0.0714}{(0.4154)(0.5462)^3} \approx 0.500$$

So $x \approx 0.0846$ M.

(c) Equilibrium concentrations: $[N_2] = 0.500 - 0.0846 = 0.4154$ M $[H_2] = 0.800 - 3(0.0846) = 0.5462$ M
$[NH_3] = 2(0.0846) = 0.1692$ M

(d) Verification:

$$K_c = \frac{(0.1692)^2}{(0.4154)(0.5462)^3} = \frac{0.02863}{(0.4154)(0.1629)} = \frac{0.02863}{0.06767} \approx 0.5008$$

The value is approximately 0.500, which matches the given K_c.