1 Explain what you understand by the term *dynamic equilibrium*.

All concentrations remain constant/macroscopic properties constant
Rate of forward reaction = rate of reverse reaction

2 Explain what is meant by a *closed system*.

No exchange of matter with surroundings

3 Describe, on a molecular level, how equilibrium is established when a liquid is placed in a closed container.

Rate of evaporation = rate of condensation

4 Write expressions for the equilibrium constant, K_c, for

$CH_4(g) + 2H_2O(g) \rightleftharpoons CO_2(g) + 4H_2(g)$
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

$K_c = \frac{[CO_2(g)][H_2(g)]^4}{[CH_4(g)][H_2O(g)]^2}$
$K_c = \frac{[NH_3(g)]^2}{[N_2(g)][H_2(g)]^3}$
$K_c = \frac{[SO_3(g)]^2}{[SO_2(g)]^2[O_2(g)]}$

5 Write expressions for the equilibrium constant, K_{c1}, for the reactions shown and state the relationship between the values of K_{c1} and K_c in 4

$CO_2(g) + 4H_2(g) \rightleftharpoons CH_4(g) + 2H_2O(g)$
$^1/_2N_2(g) + {}^3/_2H_2(g) \rightleftharpoons NH_3(g)$
$SO_3(g) \rightleftharpoons SO_2(g) + {}^1/_2O_2(g)$

$K_c = \frac{[CO_2(g)][H_2(g)]^4}{[CH_4(g)][H_2O(g)]^2}$	$K_{c1} = \frac{[CH_4(g)][H_2O(g)]^2}{[CO_2(g)][H_2(g)]^4}$	$K_{c1} = 1/K_c$
$K_c = \frac{[NH_3(g)]^2}{[N_2(g)][H_2(g)]^3}$	$K_{c1} = \frac{[NH_3(g)]}{[N_2(g)]^{1/2}[H_2(g)]^{3/2}}$	$K_{c1} = \sqrt{K_c}$
$K_c = \frac{[SO_3(g)]^2}{[SO_2(g)]^2[O_2(g)]}$	$K_{c1} = \frac{[SO_2(g)][O_2(g)]^{1/2}}{[SO_3(g)]}$	$K_{c1} = 1/\sqrt{K_c}$

6 Explain the connection between the value of the equilibrium constant and the position of equilibrium.

Bigger value of K_c – more the position of equilibrium lies towards the right
$K_c >> 1$ position of equilibrium long way to right
$K_c << 1$ position of equilibrium long way to left

7 Predict, for each of the reactions shown, the effect of the changes below on the position of equilibrium and the value of the equilibrium constant

$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$	$\Delta H = +206\ kJ\ mol^{-1}$
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$	$\Delta H = -92\ kJ\ mol^{-1}$
$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$	$\Delta H = +52\ kJ\ mol^{-1}$

increasing the pressure   decreasing the temperature   adding hydrogen   adding a catalyst

$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$	**Increase P**	**Decrease T**	**Add H_2**	**Catalyst**
Effect on position of equilibrium	To left	To left	To left	No change
Effect on K_c	No change	decreases	No change	No change

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$	**Increase P**	**Decrease T**	**Add H_2**	**Catalyst**
Effect on position of equilibrium	To right	To right	To right	No change
Effect on K_c	No change	increases	No change	No change

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$	**Increase P**	**Decrease T**	**Add H_2**	**Catalyst**
Effect on position of equilibrium	No change	To left	To right	No change
Effect on K_c	No change	decreases	No change	No change

8 Explain whether the reaction A ⇌ B is endothermic or exothermic from the values of the equilibrium constant.

$K_c = 1.2 \times 10^4$ at 500 K and $K_c = 5.6 \times 10^5$ at 1000 K

Endothermic

9 Calculate K_c for the each of the following reactions:

(a) 2A(g) + 2X(g) ⇌ 2Q(g) + Z(g) ***2250***

Initial number of moles of A = 0.200 mol	initial number of moles of X = 0.200 mol
initial number of moles of Q = 0.200 mol	initial number of moles of Z = 0.200 mol
number of moles of X at equilibrium = 0.100 mol	Volume of container = 10.0 dm^3

(b) 2A(g) + X(g) ⇌ 4Q(g) + Z(g) ***2.37×10^{-3}***

Initial number of moles of A = 0.800 mol	initial number of moles of X = 0.400 mol
initial number of moles of Q = 0.000 mol	initial number of moles of Z = 0.000 mol
number of moles of Z at equilibrium = 0.100 mol	Volume of container = 10.0 dm^3

10 Write expressions for the reaction quotient, Q, for the following reactions

2A(g) + Q(g) ⇌ 2X(g) + 3D(g)	$Q = \frac{[X]^2[D]^3}{[A]^2[Q]}$
A(g) + 3Q(g) ⇌ 4X(g) + D(g)	$Q = \frac{[X]^4[D]}{[A][Q]^3}$

11 Given the data shown and the value of K_c work out the value of Q and whether

A the system is at equilibrium
B proceeds to the right towards equilibrium
C proceeds to the left towards equilibrium

(a) A(g) + X(g) ⇌ 2Q(g) + Z(g)

Q=0.0200 therefore to the right

moles of A = 0.200 mol	moles of X = 0.200 mol
moles of Q = 0.200 mol	moles of Z = 0.200 mol
Value of K_c = 0.0300	container volume = 10.0 dm^3

(b) 2A(g) + X(g) ⇌ 2Q(g) + 2Z(g)

Q=0.0400 therefore to the left

moles of A = 0.100 mol	moles of X = 0.400 mol
moles of Q = 0.200 mol	moles of Z = 0.200 mol
Value of K_c = 0.0200	container volume = 10.0 dm^3

(c) A(g) + 2X(g) ⇌ Q(g) + 3Z(g)

Q=0.160 therefore at equilibrium

moles of A = 0.100 mol	moles of X = 0.200 mol
moles of Q = 0.100 mol	moles of Z = 0.400 mol
Value of K_c = 0.160	container volume = 10.0 dm^3

12 Calculate the number of moles present at equilibrium given the value of the equilibrium constant.

$$A_{(g)} + Z_{(g)} \rightleftharpoons X_{(g)} + Q_{(g)}$$

At a certain temperature, 0.100 mol of A and 0.100 mol of Z are placed in a container of volume 1.00 dm^3 and allowed to come to equilibrium. The value of the equilibrium constant at this temperature is 9.00. Calculate the number of moles of X at equilibrium.

0.0750 mol

13 Use Q and K_c to explain Le Chatelier's principle for changes in concentration.

$$A(g) + X(g) \rightleftharpoons Q(g) + Z(g)$$

The system at equilibrium contains the following number of moles in a vessel of volume 1.00 dm^3.

Moles of A / mol	Moles of X / mol	Moles of Q / mol	Moles of Z / mol
0.100	0.100	0.200	0.200

Explain what happens (using the values of Q and K_c) when another 0.100 mol of A is introduced into the container.

$K_c = 4$ when 0.100 mol A introduced Q = 2
$Q < K_c$ therefore the reaction is not at equilibrium and must proceed to right to reach equilibrium. New equilibrium will contain more than 0.200 mol of Q and Z, therefore the position of equilibrium has shifted to the right.

14 State how entropy and Gibbs free energy vary as a system moves towards equilibrium.

Entropy increases – maximum at equilibrium
Gibbs free energy decreases – minimum at equilibrium

15 For a system at equilibrium: $A \rightleftharpoons B$
state the relationship between the Gibbs free energy of A and B and hence the value of ΔG.

At equilibrium Gibbs free energy of A and B equal so $\Delta G=0$

16 For each of the following situations state whether the value of the equilibrium constant is less than or greater than 1.

ΔG is negative — $K_c > 1$
ΔG is positive — $K_c < 1$

17 Calculate the value of ΔG (with units) from the following data. $R=8.31\ JK^{-1}mol^{-1}$

Value of K	temperature	ΔG / kJ mol^{-1}
1.5×10^3	500 °C	-47.0
2.0×10^{-4}	300 K	21.2
10	400 K	-7.65

18 Calculate the value of the equilibrium constant from the following data $R=8.31\ JK^{-1}mol^{-1}$

ΔG/ kJ mol^{-1}	temperature	K
-100	300 °C	1.32×10^9
40	400 K	5.94×10^{-6}
-500	500 K	1.83×10^{52}

19 Explain for each of the system in 17 and 18 whether the position of equilibrium lies more to the left or more to the right.

If ΔG –ve then position of equilibrium more to right. If K > 1 then position of equilibrium to right.