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EQUILIBRIUM

K_p

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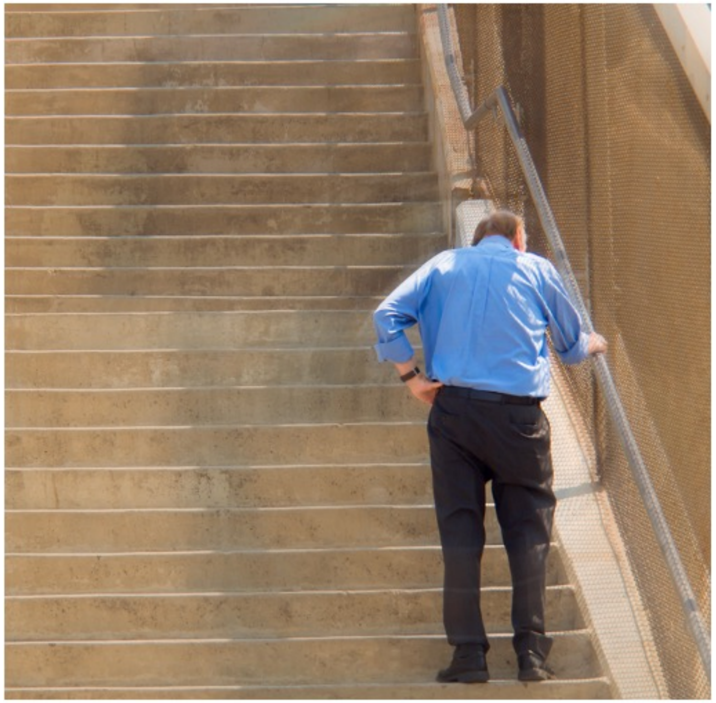
static equilibrium



person standing still on stairs

nothing is changing

at **equilibrium**

but **STATIC**

dynamic equilibrium

person walking up a down escalator
(at same speed as escalator)

nothing is changing

at **equilibrium**

but **DYNAMIC**

DYNAMIC EQUILIBRIUM

$$P + Q \rightleftharpoons R$$

- Reversible reactions can reach a state of dynamic equilibrium in a closed system
- Forward and reverse reactions taking simultaneously and at the same rate
- The concentration of all reactants and products remains constant

DYNAMIC EQUILIBRIUM

$$P + Q \rightleftharpoons R$$

- If an equilibrium **lies to the left**:
  - there are more of the substances on the left side (P&Q) than those on the right side (R)

- If an equilibrium **moves to the left**:
  - more of the backwards reaction (relative to the forwards reaction) takes place until a new equilibrium is reached, meaning once this new equilibrium is reached the amount of the substances on the left (P&Q) have increased and the amount of the substances on the right (R) have decreased

EQUILIBRIUM CONSTANT K_c

- K_c is the equilibrium constant for a system at equilibrium which uses the concentrations (in mol dm^{-3}) of the reactants and products.
- K_c is a constant for an equilibrium. The only factor that changes K_c is temperature.

$$3A + B \rightleftharpoons 2C$$

$$K_c = \frac{[C]^2}{[A]^3\,[B]}$$

EQUILIBRIUM CONSTANT K_c

Calculate K_c at temperature T for an equilibrium mixture containing 1.2 mol of A, 3.0 mol of B and 1.8 mol of C in a container with volume 2.0 dm^3

$$3A + B \rightleftharpoons 2C$$

$$K_c = \frac{[C]^2}{[A]^3\,[B]} = \frac{[\frac{1.8}{2.0}]^2}{[\frac{1.2}{2.0}]^3\,[\frac{3.0}{2.0}]} = 2.5\ mol^{-2}\,dm^6$$

$$\frac{(mol\ dm^{-3})^2}{(mol\ dm^{-3})^4} = (mol\ dm^{-3})^{-2} = mol^{-2}\,dm^6$$

LE CHATELIER'S PRINCIPLE

In simple terms – if the conditions of a system in equilibrium are changed, then the position of the equilibrium moves to oppose that change.

Gas pressure *(if a reaction involves one or more gases)*

if P ↑

equilibrium moves in the direction of the side with fewer gas molecules to lower P

equilibrium moves, but K_c/K_p remains constant.

LE CHATELIER'S PRINCIPLE

In simple terms – if the conditions of a system in equilibrium are changed, then the position of the equilibrium moves to oppose that change.

Temperature

if T ↑

equilibrium moves in the endothermic direction to lower T

equilibrium moves, and K_c/K_p change

LE CHATELIER'S PRINCIPLE

In simple terms – if the conditions of a system in equilibrium are changed, then the position of the equilibrium moves to oppose that change.

Concentration

if more of a chemical is added

equilibrium moves to use up that chemical in reaction

equilibrium moves, but K_c/K_p remains constant

K_p

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

$$K_p = \frac{(pNH_3)^2}{(pN_2)\,(pH_2)^3}$$

$$\frac{kPa^2}{kPa^4} = kPa^{-2}$$

MOLE FRACTION

If you have 10 moles of a mixture of gases, of which 3 moles is O_2, then the mole fraction of O_2 in the mixture is $\frac{3}{10}$ or 0.3

(or even 30%, but it is not usually expressed as a %)

$$\text{mole fraction of gas A in a mixture of gases} = \frac{\text{moles of gas A}}{\text{total moles of gas in the mixture}}$$

Note that the sum of all the mole fractions in a mixture should add up to 1

PARTIAL PRESSURE

Imagine a mixture of gases with a total pressure of 100 kPa.

If the mole fraction of gas A in that mixture is $\frac{3}{10}$, then that gas makes up 30 kPa (i.e $\frac{3}{10}$) of the total pressure.

The contribution that each gas makes to the total pressure is called the partial pressure of that gas; therefore the partial pressure of gas A is 30 kPa.

partial pressure of gas A in a mixture of gases = mole fraction of gas A x total pressure

Note that the sum of all the partial pressures is the total pressure.

PARTIAL PRESSURE

	$CH_4(g)$ +	$H_2O(g)$ ⇌	$CO(g)$ +	$3\ H_2(g)$
initial moles	2.00	2.00	0	0
change in moles	–1.20	–1.20	+1.20	+3.60
equilibrium moles	0.80	0.80	1.20	3.60
mole fraction	$\frac{0.80}{6.40} = 0.125$	$\frac{0.80}{6.40} = 0.125$	$\frac{1.20}{6.40} = 0.1875$	$\frac{3.60}{6.40} = 0.5625$
partial pressure	0.125 P = 31.9 kPa	0.125 P = 31.9 kPa	0.1875 P = 47.8 kPa	0.5625 P = 143.4 kPa

total moles = 6.40

total mole fractions = 1.00

total partial pressures = 255 kPa

K_p

a) At equilibrium, there are 2.0 moles of A, 2.5 moles of B and 1.5 moles of C. The total pressure is 500 kPa. Calculate the value of K_p

$$2A(g) + B(g) \rightleftharpoons 2C(g)$$

	A	B	C	
equilibrium moles	2.0	2.5	1.5	total = 6.0
mole fraction	$\frac{2.0}{6.0}$	$\frac{2.5}{6.0}$	$\frac{1.5}{6.0}$	total = 1.0
partial pressure	$\frac{2.0}{6.0}\ x\ 500$ = 166.7 kPa	$\frac{2.5}{6.0}\ x\ 500$ = 208.3 kPa	$\frac{1.5}{6.0}\ x\ 500$ = 125 kPa	total = 500

$$K_p = \frac{(pC)^2}{(pA)^2(pB)} = \frac{125^2}{166.7^2\ x\ 208.3} = 0.0027\ kPa^{-1}$$

$$\frac{kPa^2}{kPa^3} = kPa^{-1}$$

K_p

b) If the total pressure of this equilibrium was increased, explain what would happen to the yield of C and the value of K_p

$$2A(g) + B(g) \rightleftharpoons 2C(g)$$

- equilibrium shifts right
- to side with fewer gas molecules
- to lower the pressure
- so yield of C increases
- no change to K_p as it is a constant only affected by temperature

K_p

c) The forward reaction is exothermic. If the temperature of this equilibrium was increased, explain what would happen to the yield of C and the value of K_p

$$2A(g) + B(g) \rightleftharpoons 2C(g)$$

- equilibrium shifts left
- in endothermic direction
- to lower the temperature
- so yield of C decreases
- K_p decreases