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Chemical Equilibrium:
K, Q & Le Chatelier

Printable Flashcards — Pre-Med Chemistry

Equilibrium basics, K expressions, Kc vs Kp, heterogeneous equilibria, reaction quotient Q, Le Chatelier's principle (concentration, pressure, temperature, catalysts), manipulating K, ICE tables, and industrial processes.

117 cards — Print double-sided, flip on long edge, then cut along dashed lines.

117 cards — Printable Flashcards

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Generated February 20, 2026

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1

Chemical equilibrium: what's the actual meaning (not vibes)?

2

Trap: at equilibrium, the reaction stops. True or false?

3

At equilibrium, are reactant and product concentrations equal?

4

What decides whether equilibrium favors products or reactants?

5

Trap: big K means the reaction is fast. True or false?

6

K around 1 means...

7

If K is huge (like 10^{10}), what does that tell you?

8

If K is tiny (like 10^-8), what does that tell you?

2

False.

1

Forward rate = reverse rate, so concentrations stop changing (but reaction is still happening).

4

The equilibrium constant K (at that temperature).

3

Not necessarily.

6

Neither side is strongly favored; you'll have a noticeable mix of reactants and products.

5

False.

8

Equilibrium lies far to the reactants side (mostly reactants).

7

Equilibrium lies very far to the products side (mostly products).

9

At equilibrium: $\text{rate}_{\text{forward}}$ {{c1::=}} $\text{rate}_{\text{reverse}}$, so concentrations are {{c2::constant}}.

10

At equilibrium, the forward rate is ______ the reverse rate.

11

For aA + bB <-> cC + dD, Kc is...

12

Trap: you can write K using an unbalanced equation. True or false?

13

Big rule: which species do you leave OUT of K?

14

Trap: you include a solid in K because it has moles. True or false?

15

Example: CaCO3(s) <-> CaO(s) + CO2(g). Kp is...

16

Example: H2(g) + I2(g) <-> 2HI(g). Kp is...

10

equal to

9

At equilibrium: $\text{rate}_{\text{forward}} = \text{rate}_{\text{reverse}}$,
so concentrations are constant.

12

False.

11

$Kc = [C]^c [D]^d / ([A]^a [B]^b)$.

14

False.

13

Pure solids (s) and pure liquids (l).

16

$Kp = (P_H I)^2 / (P_{H2} * P_{I2})$.

15

Kp = P(CO2).

17

Kp uses ____, Kc uses ____

18

Trap: Kp and Kc are always equal. True or false?

19

Δn_{gas} (for Kp vs Kc) means...

20

Relationship between Kp and Kc:

21

Heterogeneous equilibrium rule: omit pure {{c1::solids}} and pure {{c2::liquids}} from K.

22

In K expressions, do you include pure liquids (l)?

23

Reaction quotient Q is basically...

24

How do you predict direction using Q and K?

18

False.

17

Kp uses partial pressures (gases).
Kc uses concentrations (mol/L).

20

$Kp = Kc\ (RT)\hat{}(\Delta n_{gas})$.

19

Moles of gaseous products minus moles of gaseous reactants.

22

No

21

Heterogeneous equilibrium rule: omit pure solids and pure liquids from K.

24

If $Q < K$ -> shift forward (make products). If $Q > K$ -> shift backward (make reactants). If $Q = K$ -> already at equilibrium.

23

The same expression as K, but using CURRENT concentrations/pressures (not equilibrium ones).

25

Trap: if Q is bigger than K, reaction speeds up forward. True or false?

26

At equilibrium, Q equals...

27

$N_2 + 3H_2 \leftrightarrow 2NH_3$. If you suddenly add NH_3, what happens to Q and which way does it shift?

28

Same reaction: if you remove H_2, what happens to Q and shift direction?

29

Same reaction: if you add N_2, what happens to Q and shift direction?

30

Fast rule (memorize): adding reactant shifts...

31

Fast rule: adding product shifts...

32

Fast rule: removing product shifts...

26

K.

25

False.

28

Q increases, so it shifts left.

27

Q increases, so it shifts left (toward reactants).

30

Right (toward products).

29

Q decreases, so it shifts right (toward products).

32

Right.

31

Left (toward reactants).

33

Fast rule: removing reactant shifts...

34

Compare Q and K: if Q {{c1::<}} K, shift {{c2::right}}; if Q {{c3::>}} K, shift {{c4::left}}.

35

If $Q < K$, the reaction shifts to the ______.

36

If you change concentrations at constant temperature, does K change?

37

Trap: adding more reactant increases K. True or false?

38

For A(aq) <-> B(aq) + C(aq), what happens if you dilute the solution (add water)?

39

Trap: dilution always shifts right. True or false?

40

Heterogeneous equilibrium trap: CaCO3(s) <-> CaO(s) + CO2(g). If you add more CaCO3(s), what happens to equilibrium CO2 pressure?

34

Compare Q and K: if $Q < K$, shift right; if $Q > K$, shift left.

33

Left.

36

No.

35

right

38

It shifts to the side with MORE dissolved particles (here: to the right).

37

False.

40

Nothing (no change), as long as solids are still present.

39

False.

41

Same reaction: if you remove some CaCO3(s) but still leave some solid behind, what happens?

42

But if you remove ALL of one solid in a heterogeneous equilibrium, what happens?

43

At constant temperature, changing concentration changes {{c1::Q}}, not {{c2::K}}.

44

Pressure changes only affect equilibria involving...

45

When volume decreases (pressure increases), equilibrium shifts to the side with...

46

When volume increases (pressure decreases), equilibrium shifts to the side with...

47

Trap: pressure change affects both sides even if gas moles are equal. True or false?

48

Example: N2(g) + 3H2(g) <-> 2NH3(g). Decreasing volume shifts...

42

The system can't maintain the same equilibrium; the reaction stops shifting once that phase is gone.

41

Still no change to equilibrium CO2 pressure.

44

Gases.

43

At constant temperature, changing concentration changes Q, not K.

46

More moles of gas.

45

Fewer moles of gas.

48

Right (toward NH3).

47

False.

49

Example: $N_2O_4(g)$ <-> $2NO_2(g)$. Increasing pressure shifts...

50

Pressure trick: instead of guessing, you can do this clean method:

51

Inert gas trap (constant V): If you add an inert gas to a sealed rigid container, does equilibrium shift?

52

Inert gas (constant P): If you add an inert gas but keep total pressure constant (so volume increases), equilibrium shifts to the side with...

53

Trap: adding inert gas ALWAYS shifts equilibrium. True or false?

54

Decrease volume (increase pressure) -> shift to {{c1::fewer}} moles of gas. Increase volume -> shift to {{c2::more}} moles of gas.

55

If pressure increases, equilibrium shifts toward the side with ______ moles of gas.

56

Key fact: K changes only when you change...

50

Compute Q after the change (new partial pressures) and compare to K.

49

Left (toward N2O4).

52

More moles of gas.

51

No.

54

Decrease volume (increase pressure) -> shift to fewer moles of gas. Increase volume -> shift to more moles of gas.

53

False.

56

Temperature.

55

fewer

57

Le Chatelier for temperature: treat heat as...

58

If the forward reaction is exothermic (releases heat), increasing temperature shifts equilibrium...

59

If the forward reaction is endothermic (absorbs heat), increasing temperature shifts equilibrium...

60

How does K change when temperature increases for an endothermic forward reaction?

61

How does K change when temperature increases for an exothermic forward reaction?

62

Trap: catalysts change K because they lower activation energy. True or false?

63

Haber process ($N_2 + 3H_2 \leftrightarrow 2NH_3$) is exothermic. For more NH_3, temperature should be...

64

Color trick: $N_2O_4(g) \leftrightarrow 2NO_2(g)$. NO_2 is brown, N_2O_4 is colorless. Heating shifts toward...

58

Left (toward reactants).

57

A reactant in endothermic reactions,
and a product in exothermic reactions.

60

K increases.

59

Right (toward products).

62

False.

61

K decreases.

64

NO2 (browner).

63

Lower (but then the reaction is
slower, so we use a catalyst).

65

K only changes with {{c1::temperature}}. (Not with concentration, pressure, or catalysts.)

66

For an exothermic forward reaction, increasing temperature shifts equilibrium to the ______.

67

What does a catalyst actually do to an equilibrium system?

68

Trap: catalyst shifts equilibrium toward products. True or false?

69

K tells you about ______ ,
rate tells you about ______

70

A reaction can be product-favored (big K) but still take forever because...

71

Trap: at equilibrium, forward and reverse rates are both zero. True or false?

72

If you reverse a reaction, what happens to K?

66

left

65

K only changes with temperature. (Not with concentration, pressure, or catalysts.)

68

False.

67

It makes equilibrium happen faster (speeds up forward and reverse rates equally).

70

It has a big activation energy (slow kinetics).

69

K tells you how far the reaction goes (equilibrium position). Rate tells you how fast it gets there.

72

K becomes 1/K.

71

False.

73

If you multiply the entire balanced equation by n, what happens to K?

74

If you ADD two reactions to get a new overall reaction, what happens to K overall?

75

If you SUBTRACT a reaction (or reverse it then add), what happens to K?

76

Trap: if you change coefficients, K stays the same because equilibrium is 'the same'. True or false?

77

Example: If K for A <-> B is 4, what is K for B <-> A?

78

Example: If K for A <-> B is 4, what is K for 2A <-> 2B?

79

Reverse reaction: K_{new} = {{c1::1/K}}.
Multiply reaction by n: K_{new} = {{c2::K^n}}.

80

If $K_{\text{forward}} = 10$, $K_{\text{reverse}} =$

74

Multiply the K values.

73

K becomes K^n.

76

False.

75

You divide (or multiply by $1/K$) accordingly.

78

$K^2 = 16$.

77

$1/4 = 0.25$.

80

0.1

79

Reverse reaction: $K_{new} = 1/K$.
Multiply reaction by n: $K_{new} = K^n$.

81

ICE table stands for...

82

When do you actually need an ICE table?

83

Common ICE trap: students use coefficients in the 'Change' row wrong. Fix it:

84

Example pattern: A <-> 2B, start with [A]=1.0, [B]=0.0. At equilibrium, if A decreased by x then B is...

85

pre-med level approximation trick: if K is very small and you start mostly reactants, then x is usually...

86

Trap: you can always ignore x in ICE tables. True or false?

87

ICE stands for {{c1::Initial}}, {{c2::Change}}, {{c3::Equilibrium}}.

88

Haber process: N2 + 3H2 <-> 2NH3. Pressure increase favors which side?

82

When you're given K and some starting concentrations and you must solve for unknown equilibrium concentrations.

81

Initial, Change, Equilibrium.

84

2x.

83

Changes must follow the mole ratio from the balanced equation.

86

False.

85

Small (so reactant concentration barely changes).

88

Products (NH3).

87

ICE stands for Initial, Change, Equilibrium.

89

Haber process: temperature increase favors which side (forward is exothermic)?

90

Industry compromise for Haber (concept): they use...

91

Contact process step: 2SO2(g) + O2(g) <-> 2SO3(g). Pressure increase shifts...

92

N2O4(g) <-> 2NO2(g). Volume decrease (pressure increase) shifts...

93

Same system: N2O4 is colorless, NO2 is brown. Cooling the system makes it...

94

CaCO3(s) <-> CaO(s) + CO2(g). If you increase temperature (decomposition is endothermic), CO2 pressure at equilibrium...

95

For N2O4(g) <-> 2NO2(g), increasing pressure shifts equilibrium to the:

96

Quick sorting (pre-med favorite): changing concentration affects...

90

High pressure, moderate temperature, and a catalyst.

89

Reactants.

92

Left (toward N2O4).

91

Right (toward SO3).

94

Increases (more CO2 formed).

93

Less brown (more N2O4).

96

Q (and equilibrium position), not K.

95

left

97

Changing pressure/volume affects...

98

Changing temperature affects...

99

Catalyst affects...

100

Trap: if you remove product continuously, K changes because the system keeps shifting. True or false?

101

Trap: if you increase pressure, Kp changes because pressures are in Kp. True or false?

102

Only {{c1::temperature}} changes K. Concentration/pressure changes affect {{c2::Q}} and shift equilibrium.

103

A <-> B with K = [B]/[A]. If you double BOTH [A] and [B] instantly, does Q change?

104

A + B <-> C with K = [C]/([A][B]). If you double all concentrations instantly, Q becomes...

98

K (and therefore equilibrium position).

97

Q (for gases) and equilibrium position,
not K (if temperature stays the same).

100

False.

99

Rates (time to reach equilibrium),
not K or equilibrium position.

102

Only temperature changes K.
Concentration/pressure changes
affect Q and shift equilibrium.

101

False.

104

Half of what it was (Q decreases).

103

No (Q stays the same).

105

For 2A <-> B, $K = [B]/[A]^2$. If [A] doubles (B constant), Q...

106

Fast check: if you add reactant, Q goes...

107

Fast check: if you add product, Q goes...

108

Thermo link (if they mention it): $\Delta G°$ is related to K by...

109

If $K > 1$, $\Delta G°$ is...

110

If $K < 1$, $\Delta G°$ is...

111

Trap: $\Delta G° = 0$ means no reaction happens. True or false?

112

If K is very large and you start with reactants only, what will Q vs K look like at the start?

106

Down.

105

Decreases by a factor of 4.

108

$\Delta G^\circ = -RT \ln K$.

107

Up.

110

Positive.

109

Negative.

112

Q starts near 0, so $Q < K$ -> shifts right strongly.

111

False.

113

If you start with only products
for a reversible reaction, Q is...

114

You are given an equilibrium with no gases involved. If pressure changes, what happens?

115

If a reaction has equal moles of gas on both sides, compressing the container causes...

116

Trap: 'shift right' means K increases. True or false?

117

The only factor that changes
the numerical value of K is:

114

Nothing (for equilibrium position).

113

Very large (often 'infinite' if reactants are 0), so it shifts left.

116

False.

115

No shift in equilibrium position.

117

temperature