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Topic 7.1: Introduction to Equilibrium

Worksheet · AP Chemistry

UNIT 7 · EQUILIBRIUM

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How to Use This Worksheet

Work in order. Section A checks the foundations: reversible reactions, the conditions equilibrium needs, and how the two rates behave on the way there. Section B works with one data set, $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$, from a concentration table to rates, a second flask, pressure and a disturbance. Section C is exam style: the N_2O_4 and NO_2 color system, a rate-time graph, and six claims to judge.

Reference facts. Equilibrium means equal forward and reverse rates, so every concentration stays constant; the rate ratio is exactly 1. The system must be closed to matter and held at constant temperature. The same equilibrium is reached from either side with the same totals. Useful values: the SO_2 system settles at $[SO_2] = 0.40$, $[O_2] = 0.30$ and $[SO_3] = 0.60$ mol L^{-1}.

Three habits protect every answer. Keep the two conditions of equilibrium separate: equal rates, and constant concentrations; most wrong answers swap one for the other. Read graphs for plateaus and meeting points, not crossings. Check the system is closed before applying equilibrium reasoning.

Do not look at the answer key until you have attempted every question.

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Section A: Recall and Concepts

A1. (a) State what is meant by a reversible reaction. (b) State what the symbol ⇌ tells you that a single arrow → does not. (c) State the two conditions needed for equilibrium to be established: one about the system itself, one about its temperature.

A2. (a) State the condition on the forward and reverse rates at equilibrium. (b) State the condition on the concentrations at equilibrium. (c) Explain why these are two different statements, not one statement said twice.

A3. A reversible reaction is started with reactants only in a sealed flask. Describe how each of the following changes from the moment of mixing until equilibrium is reached, giving a reason for each: (a) the forward rate; (b) the reverse rate; (c) the reactant concentration; (d) the product concentration.

A4. (a) On a concentration-time graph, state the feature that shows equilibrium has been reached. (b) On a rate-time graph, state the feature that shows equilibrium has been reached. (c) At the moment the two rate curves meet, state the value of (forward rate) ÷ (reverse rate). (d) State whether the shared rate after the curves meet is zero, and explain your answer.

A5. A flask of $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ is at equilibrium. A trace of I_2 labelled with a radioactive isotope is added, far too little to change any concentration measurably. (a) State where the label is found as time passes. (b) State what this proves about the two reactions at equilibrium. (c) Explain why the bulk composition does not change while the label spreads.

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Section B: Core Practice

The reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ is followed in a sealed vessel at constant temperature. Concentrations in mol L^{-1}:

Time (min)	$[SO_2]$	$[O_2]$	$[SO_3]$
0	1.00	0.60	0.00
2	0.70	0.45	0.30
4	0.52	0.36	0.48
6	0.44	0.32	0.56
8	0.40	0.30	0.60
10	0.40	0.30	0.60

B1. (a) State the time by which equilibrium has been reached, and the evidence in the table. (b) State the equilibrium concentration of each species. (c) Calculate the change in concentration of each species between 0 and 8 minutes.

B2. (a) Show that the changes found in B1(c) obey the stoichiometry of the equation. (b) Calculate the average rate of formation of SO_3 between 0 and 2 minutes. (c) Calculate the average rate of formation of SO_3 between 8 and 10 minutes, and explain why it has that value even though SO_3 is still reacting.

B3. (a) State the reverse rate at time zero, and explain why it has that value. (b) Explain, in terms of concentrations, why the forward rate decreases and the reverse rate increases during the first 8 minutes. (c) At equilibrium the forward and reverse rates are each 0.018 mol L^{-1} min^{-1}. A student says the net rate of change of $[SO_3]$ must therefore be 0.018 mol L^{-1} min^{-1}. Correct the student.

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B4. A second vessel at the same temperature is charged with $[SO_3]$ = 1.00 mol L^{-1} and $[O_2]$ = 0.10 mol L^{-1}, and no SO_2. (a) State which reaction, forward or reverse as the equation is written in B1, runs faster at the start, and why. (b) Show that this vessel holds the same total amount of sulfur and of oxygen per litre as the vessel in B1. (c) Predict the equilibrium concentrations in this vessel, and justify the prediction. (d) State the change in $[SO_3]$ needed to reach that equilibrium.

B5. (a) Calculate the total concentration of gas at time zero and at equilibrium in the B1 vessel. (b) At constant volume and temperature, pressure is proportional to total concentration. Calculate the ratio (pressure at equilibrium) ÷ (initial pressure). (c) Explain why the pressure in this vessel falls as the reaction proceeds, in terms of the equation. (d) State what a pressure gauge would show at 6 minutes and at 20 minutes, and what each reading tells you.

B6. The B1 vessel is at equilibrium when extra SO_2 is injected, raising $[SO_2]$ suddenly. (a) State what happens to the forward rate immediately, and why. (b) State what happens to the reverse rate at the instant of injection, and afterwards as product builds up. (c) Describe the new equilibrium: which concentrations are higher than in the original equilibrium, which is lower, and which rate condition holds again. (d) Explain why the system does not simply remain at the original equilibrium.

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Section C: Exam Style

C1. NO_2 is a brown gas and N_2O_4 is colorless: $N_2O_4(g) \rightleftharpoons 2NO_2(g)$. Tube 1 is sealed with $[N_2O_4] = 0.50$ mol L^{-1} and no NO_2. The brown color deepens steadily, then stops deepening. At equilibrium $[N_2O_4] = 0.10$ mol L^{-1} and $[NO_2] = 0.80$ mol L^{-1}. (a) Explain the color behavior at the particulate level, including why the color stops deepening while both gases are still reacting. (b) Calculate the amount of N_2O_4 per litre that has reacted by equilibrium in tube 1, and check it against the NO_2 formed. (c) Tube 2, at the same temperature, is sealed with $[NO_2] = 1.00$ mol L^{-1} only. Predict its equilibrium concentrations, justifying your answer with the total amounts of nitrogen and oxygen in each tube. (d) A student writes: "At equilibrium, tube 1 must contain equal concentrations of the two gases, because the reactions have balanced." Judge the statement.

C2. A rate-time graph is drawn for $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ in a sealed vessel. The forward rate starts at 0.48 mol L^{-1} min^{-1} and falls steadily. The reverse rate starts at zero and rises. The two curves meet at 0.09 mol L^{-1} min^{-1} at t = 12 min and run level together after that. (a) State when equilibrium is reached, and the evidence. (b) At t = 4 min, state which rate is larger, and what is happening to $[PCl_5]$ then. (c) Explain why the forward rate falls and the reverse rate rises during the first 12 minutes. (d) The experiment is repeated with a catalyst present from the start. State the effect on the time taken to reach equilibrium, on the value at which the curves meet, and on the equilibrium composition, explaining each.

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C3. Judge each claim, supporting each verdict with the idea it depends on. (a) "At equilibrium the reaction has stopped, which is why the concentrations no longer change." (b) "Equilibrium is reached when the concentration of reactants equals the concentration of products." (c) "A reversible reaction in an open beaker reaches the same equilibrium as in a sealed flask." (d) "If the forward rate is greater than the reverse rate, the system cannot be at equilibrium, even if the concentrations have looked steady for a minute." (e) "Adding a catalyst moves the equilibrium position toward the products." (f) "Equilibrium can only be approached from the reactants side; starting from pure products just runs the reaction backwards to completion."

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Answer Key

A1. (a) A reaction in which the products can react together to re-form the reactants, so both directions occur in the same mixture. (b) The double arrow states that both directions happen and neither goes to completion; a single arrow treats the reaction as running one way to completion. (c) The system must be closed to matter, and the temperature constant, because the rates and the settling composition depend on temperature.

A2. (a) The forward rate equals the reverse rate. (b) Every concentration stays constant over time. (c) The rate condition is about how fast particles convert in each direction; the concentration condition is about the amounts present. Equal rates produce constant concentrations, but the concentrations need not be equal, so one statement cannot replace the other.

A3. (a) The forward rate starts at its maximum, because reactant concentration is at its highest, and decreases as reactants are used up. (b) The reverse rate starts at zero, because there is no product to react back, and increases as product concentration builds. (c) Reactant concentration falls, quickly at first and then more slowly, as the net rate narrows toward zero. (d) Product concentration rises from zero, quickly at first then flattening, until it levels off at its equilibrium value.

A4. (a) Both concentration curves become horizontal plateaus: neither concentration changes with time. (b) The forward and reverse rate curves meet and run together at one shared, level value. (c) Exactly one, because the rates are equal. (d) No. The shared rate is positive: both reactions continue at that rate, which keeps the concentrations constant rather than frozen.

A5. (a) The label appears in HI and becomes distributed between I_2 and HI, while all three concentrations stay at their equilibrium values. (b) It proves both reactions are still running at equilibrium: labeled I_2 can only reach HI through the forward reaction, and its continued exchange shows the reverse reaction running too. (c) The two rates are equal, so every I_2 converted to HI is matched by an HI converted back; the amounts stay fixed even though individual molecules keep swapping.

B1. (a) By 8 minutes: the readings at 8 and 10 minutes are identical for all three species, so nothing is changing. (b) $[SO_2]$ = 0.40 mol L^{-1}, $[O_2]$ = 0.30 mol L^{-1}, $[SO_3]$ = 0.60 mol L^{-1}. (c) SO_2: 0.40 − 1.00 = −0.60 mol L^{-1}. O_2: 0.30 − 0.60 = −0.30 mol L^{-1}. SO_3: 0.60 − 0.00 = +0.60 mol L^{-1}.

B2. (a) The equation consumes SO_2 and O_2 and produces SO_3 in the ratio 2 : 1 : 2. The measured changes are 0.60 : 0.30 : 0.60, which is 2 : 1 : 2, so the table obeys the stoichiometry exactly. (b) (0.30 − 0.00) ÷ (2 − 0) = 0.15 mol L^{-1} min^{-1}. (c) (0.60 − 0.60) ÷ (10 − 8) = 0 mol L^{-1} min^{-1}. The average is zero because the concentration does not change: at equilibrium SO_3 is formed and consumed at equal rates, so the net change is zero.

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B3. (a) Zero. There is no SO_3 present at time zero, so the reverse reaction has nothing to react and cannot occur at all. (b) The forward rate falls because $[SO_2]$ and $[O_2]$ fall as they are consumed, so reactant particles collide less often. The reverse rate rises because $[SO_3]$ rises from zero, so product particles react back more often. Both trends continue until the rates meet at 8 minutes. (c) The net rate of change is the forward rate minus the reverse rate: $0.018 - 0.018 = 0$ mol L^{-1} min^{-1}. The value 0.018 mol L^{-1} min^{-1} is both the rate at which SO_3 is formed and the rate at which it is consumed; concentration changes only by the difference.

B4. (a) The reverse reaction. The vessel starts with SO_3 at high concentration and no SO_2, so the forward rate starts at zero and the reverse rate at its maximum. (b) Sulfur: both vessels hold 1.00 mol L^{-1} of sulfur atoms. Oxygen: vessel 1 holds 2×1.00 from SO_2 plus 2×0.60 from O_2, total 3.20 mol L^{-1}; vessel 2 holds 3×1.00 from SO_3 plus 2×0.10 from O_2, total 3.20 mol L^{-1}. (c) The same as vessel 1: $[SO_2] = 0.40$ mol L^{-1}, $[O_2] = 0.30$ mol L^{-1}, $[SO_3] = 0.60$ mol L^{-1}. Same reaction, temperature and totals must settle at the same equilibrium composition, whichever side they approach from. (d) $[SO_3]$ must fall from 1.00 to 0.60 mol L^{-1}, a change of -0.40 mol L^{-1}. Check: the reverse reaction converts that 0.40 mol L^{-1} of SO_3 into 0.40 mol L^{-1} of SO_2 and 0.20 mol L^{-1} of O_2, taking $[O_2]$ to 0.30 mol L^{-1}.

B5. (a) At time zero: $1.00 + 0.60 + 0.00 = 1.60$ mol L^{-1}. At equilibrium: $0.40 + 0.30 + 0.60 = 1.30$ mol L^{-1}. (b) $1.30 \div 1.60 = 0.8125$. The equilibrium pressure is 0.8125 times the initial pressure. (c) The equation turns 3 mol of gas ($2SO_2 + O_2$) into 2 mol of gas ($2SO_3$) per cycle as written, so the total amount of gas falls as the forward reaction proceeds, and pressure falls with it. (d) At 6 minutes the reading would still be falling, because the concentrations are still changing then. At 20 minutes the reading would be steady at the equilibrium pressure: equilibrium was reached by 8 minutes and the composition no longer changes.

B6. (a) The forward rate increases immediately, because $[SO_2]$ has just been raised, so reactant particles collide more often. (b) At the instant of injection the reverse rate is unchanged, because $[SO_3]$ has not changed yet. Afterwards, as the forward reaction builds up SO_3, the reverse rate climbs until it matches the forward rate again. (c) The system settles at a new equilibrium with $[SO_3]$ higher than 0.60 mol L^{-1}, $[SO_2]$ higher than its original 0.40 mol L^{-1} and $[O_2]$ lower than 0.30 mol L^{-1}, and the two rates equal again, at a higher shared value than before. (d) The original equilibrium was a balance at the original concentrations. Raising $[SO_2]$ unbalances the rates, so net forward reaction runs until concentrations have moved enough for the rates to match again.

C1. (a) At the start only the forward reaction can run, so NO_2 accumulates and the brown color deepens. As $[NO_2]$ grows, the reverse reaction speeds up, while the forward rate falls as $[N_2O_4]$ falls. The color stops deepening when the rates become equal: NO_2 is then formed and removed at the same rate, so its concentration and the color stay constant even though both reactions continue. (b) N_2O_4 reacted: $0.50 - 0.10 = 0.40$ mol L^{-1}. Check: each N_2O_4 forms $2NO_2$, so NO_2 formed should be $2 \times 0.40 = 0.80$ mol L^{-1}, matching the equilibrium value. (c) Both tubes hold the same totals: tube 1 has nitrogen $2 \times 0.50 = 1.00$ mol L^{-1} and oxygen $4 \times 0.50 = 2.00$ mol L^{-1}; tube 2

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has nitrogen 1.00 mol L^{-1} and oxygen 2 × 1.00 = 2.00 mol L^{-1}. Same reaction, temperature and totals, so tube 2 reaches the same equilibrium: $[N_2O_4]$ = 0.10 mol L^{-1} and $[NO_2]$ = 0.80 mol L^{-1}. In tube 2, 0.20 mol L^{-1} of NO_2 is consumed to form 0.10 mol L^{-1} of N_2O_4, leaving 0.80 mol L^{-1} of NO_2. (d) The statement is wrong. The reactions balance in rate, not in amount: the equilibrium values are 0.10 and 0.80 mol L^{-1}, constant but eightfold different. Equilibrium requires equal rates, never equal concentrations.

C2. (a) Equilibrium is reached at t = 12 min, when the two rate curves meet; after that both rates stay equal and level. (b) The forward rate is larger. The reverse rate has not yet caught up, so there is net forward reaction and $[PCl_5]$ is falling. (c) The forward rate falls because $[PCl_5]$ falls, so its particles collide less often. The reverse rate rises because $[PCl_3]$ and $[Cl_2]$ rise from zero, so product particles react back more often. Each rate is driven by the concentrations on its own side, which is why the curves move toward each other. (d) The catalyst raises both rates from the start, so the curves meet sooner: equilibrium is reached in less than 12 minutes, at a higher shared value than 0.09 mol L^{-1} min^{-1}. The equilibrium composition is unchanged, because the catalyst speeds both directions equally and the system still settles where the rates match.

C3. (a) False. The concentrations stop changing because the forward and reverse rates are equal, not because the reactions stop. Both reactions continue, as tracer experiments show. (b) False. Equilibrium requires constant concentrations, not equal ones. In the N_2O_4 system the equilibrium concentrations are 0.10 and 0.80 mol L^{-1}; in the H_2 and I_2 system, 0.22 and 1.56 mol L^{-1}. (c) False. In an open beaker a gaseous species can escape before reacting back, starving the reverse reaction. Equilibrium needs a closed system; the sealed flask can reach it and the open beaker cannot. (d) True. Equilibrium is defined by equal rates. If the forward rate exceeds the reverse rate there is net forward reaction and the concentrations are still changing, however steady they look over a short interval. (e) False. A catalyst raises both rates together, so equilibrium is reached sooner at the same composition. (f) False. Equilibrium can be approached from either side. Starting from pure products, the reverse reaction runs first and the system settles at the same composition as one started from reactants, given the same totals and temperature.

Continue Revising

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Topic 7.1 Quiz: Introduction to Equilibrium

Previous Worksheet: Topic 6.10 Enthalpy of Solution

Next Worksheet: Topic 7.2 Direction of Reversible Reactions

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