Ice Table Practice Problems With Answers

MASTERING CHEMICAL EQUILIBRIUM WITH ICE TABLE PRACTICE PROBLEMS WITH ANSWERS

Understanding chemical equilibrium is a cornerstone of general chemistry and advanced studies in physical chemistry. One of the most powerful tools for solving equilibrium problems is the ICE table—a systematic method that organizes Initial concentrations, Changes in concentrations, and Equilibrium concentrations of reactants and products. For students and professionals alike, working through **ICE Table Practice Problems With Answers** is the most effective way to build confidence and accuracy. This article provides a step-by-step guide to constructing and solving ICE tables, along with multiple practice problems of varying difficulty, each with a detailed answer. By the end, you will be able to tackle equilibrium calculations with ease.

What Is an ICE Table?

An ICE table is a structured format used to track the concentrations (or partial pressures) of chemical species as a reaction approaches equilibrium. The acronym stands for:

- **I** – Initial concentration (or pressure) of each species before any change occurs.
- **C** – Change in concentration as the reaction proceeds toward equilibrium (expressed in terms of an unknown variable, usually *x*).
- **E** – Equilibrium concentration, calculated by adding the initial and change values.

The ICE table is especially useful when you know the equilibrium constant (K_c or K_p) and need to find unknown equilibrium concentrations, or when you have equilibrium data and need to compute the constant. Using *ICE Table Practice Problems With Answers* reinforces the logic of stoichiometric coefficients and the sign conventions for change.

How to Set Up an ICE Table

1. **Write the balanced chemical equation.** Note the coefficients of all reactants and products.
2. **List all species in the first row (I).** Use initial concentrations or pressures given in the problem. If a substance is a pure solid or liquid, its concentration remains constant and is omitted from K expressions.
3. **Define the change (C) using a variable.** For every mole of reactant consumed, the change is –*x* times its coefficient; for every mole of product formed, the change is +*x* times its coefficient.
4. **Write equilibrium expressions (E).** E = I + C for each species.
5. **Plug equilibrium values into the equilibrium constant expression.** Solve for *x* (often using algebra, possibly the quadratic formula).
6. **Use *x* to find the required concentrations**

or pressures.

Now let's apply these steps to several *ICE Table Practice Problems With Answers*.

PRACTICE PROBLEM 1: SIMPLE EQUILIBRIUM WITH KNOWN K_c

Problem: Consider the reaction: $N_2(g) + 3 H_2(g) \rightleftharpoons 2 NH_3(g)$. Initially, 0.200 M N_2 and 0.600 M H_2 are placed in a flask. At equilibrium, the concentration of NH_3 is found to be 0.120 M. What is the equilibrium constant K_c?

Solution with ICE Table

Set up an ICE table (concentrations in M):

	N_2	H_2	NH_3
Initial	0.200	0.600	0
Change	–x	–3x	+2x
Equilibrium	0.200 – x	0.600 – 3x	2x

We are told $[NH_3]_{eq}$ = 0.120 M, therefore 2x = 0.120 ⇒ x = 0.060 M. Now calculate the equilibrium concentrations:

- $[N_2]$ = 0.200 – 0.060 = 0.140 M
- $[H_2]$ = 0.600 – 3(0.060) = 0.600 – 0.180 = 0.420 M
- $[NH_3]$ = 0.120 M

Write the K_c expression: $K_c = [NH_3]^2 / ([N_2][H_2]^3)$

$K_c = (0.120)^2 / (0.140 \times (0.420)^3) = 0.0144 / (0.140 \times 0.074088) = 0.0144 / 0.0103723 \approx 1.39$

Answer: $K_c \approx 1.39$

This first example shows how an ICE table can be used to find K_c when one equilibrium concentration is known. Many *ICE Table Practice Problems With Answers* follow this pattern.

PRACTICE PROBLEM 2: FINDING EQUILIBRIUM CONCENTRATIONS FROM K_c

Problem: For the reaction $H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g)$, $K_c = 54.3$ at 425°C. Initially, 0.100 M H_2 and 0.100 M I_2 are mixed. Calculate the equilibrium concentration of HI.

Solution

ICE table in M:

	H_2	I_2	HI
Initial	0.100	0.100	0
Change	–x	–x	+2x
Equilibrium	0.100 – x	0.100 – x	2x

$K_c = [HI]^2 / ([H_2][I_2]) = (2x)^2 / ((0.100 – x)(0.100 – x)) = 4x^2 / (0.100 – x)^2 = 54.3$

Take square root of both sides: $2x / (0.100 – x) = \sqrt{54.3} \approx 7.37$

$2x = 7.37(0.100 – x) \Rightarrow 2x = 0.737 – 7.37x \Rightarrow 2x + 7.37x = 0.737 \Rightarrow 9.37x = 0.737 \Rightarrow x \approx 0.0787$ M

Therefore $[HI]_{eq} = 2x = 2 \times 0.0787 \approx 0.157$ M

Answer: [HI] = 0.157 M

Notice how using the square root simplified the algebra because the initial concentrations were equal. This is a common shortcut in *ICE Table Practice Problems With Answers*.

PRACTICE PROBLEM 3: USING THE QUADRATIC FORMULA

Problem: For the dissociation of $SO_2Cl_2(g) \rightleftharpoons SO_2(g) + Cl_2(g)$, $K_c = 0.045$ at 375 K. If 0.200 M SO_2Cl_2 is placed in a container, what are the equilibrium concentrations of all species?

Solution

ICE table (concentrations in M):

	SO_2Cl_2	SO_2	Cl_2
Initial	0.200	0	0
Change	–x	+x	+x
Equilibrium	0.200 – x	x	x

$K_c = [SO_2][Cl_2] / [SO_2Cl_2] = (x)(x) / (0.200 - x) = x^2 / (0.200 - x) = 0.045$

Multiply: $x^2 = 0.045(0.200 - x) \Rightarrow x^2 = 0.0090 - 0.045x \Rightarrow x^2 + 0.045x - 0.0090 = 0$

Use quadratic formula: $x = [-b \pm \sqrt{(b^2 - 4ac)}] / 2a$ with a=1, b=0.045, c= –0.0090

Discriminant: $(0.045)^2 - 4(1)(-0.0090) = 0.002025 + 0.0360 = 0.038025$

$\sqrt{0.038025} \approx 0.1950$

$x = (-0.045 \pm 0.1950) / 2$. Positive root: $x = (0.1500) / 2 = 0.0750$ M

Therefore:

- $[SO_2Cl_2] = 0.200 - 0.0750 = 0.125$ M
- $[SO_2] = 0.0750$ M
- $[Cl_2] = 0.0750$ M

Answer: $[SO_2Cl_2] = 0.125$ M, $[SO_2] = [Cl_2] = 0.0750$ M

Quadratic equations appear frequently in *ICE Table Practice Problems With Answers*. Always check the 5% rule: if x is less than 5% of the initial concentration, you can approximate 0.200 –