Week 2 MCAT Practice Questions

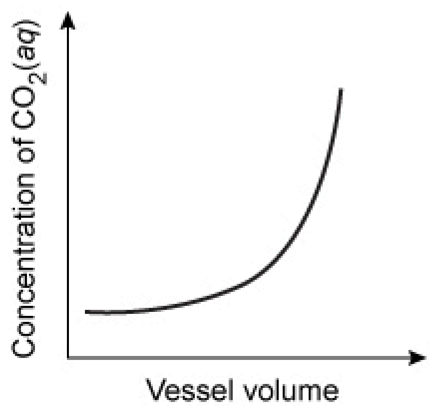
1)

A sealed reaction vessel with a movable piston contains $CO_2(g)$ in equilibrium with $H_2O(l)$ at 25 °C, as shown by the dissolution equation below.

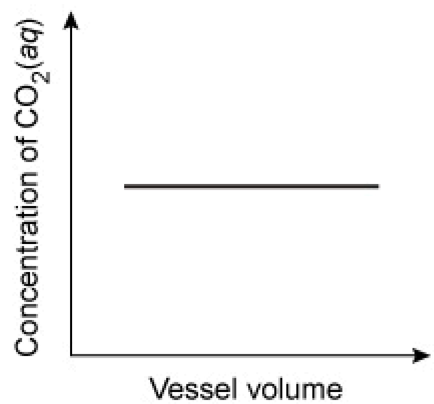
$$CO_2(g) \underset{25\ °\mathrm{C}}{\overset{\text{water}}{\rightleftharpoons}} CO_2(aq)$$

If the volume of the reaction vessel is increased, which graph best describes the effect of volume on the concentration of dissolved $CO_2(aq)$?

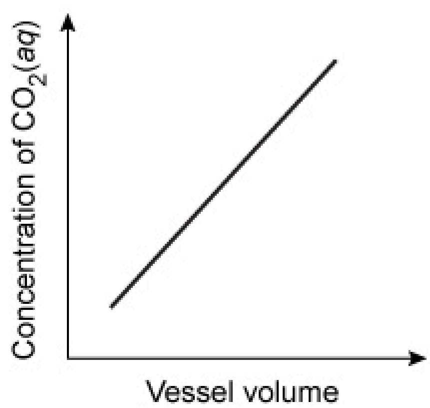
A.



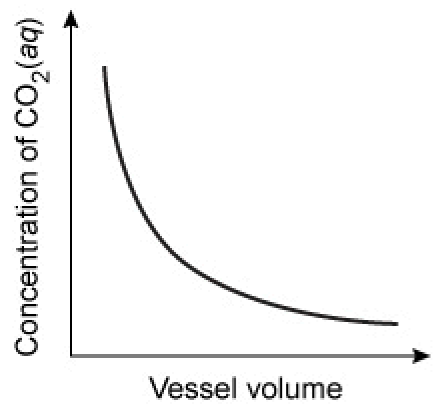
C.



B.



D.



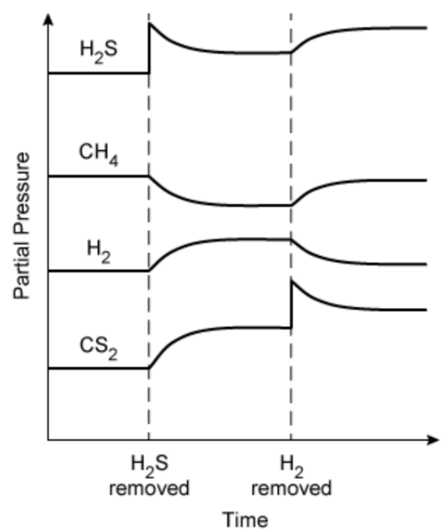
2)

A sealed, rigid container is filled with a mixture of $H_2S(g)$, $CH_4(g)$, $H_2(g)$, and $CS_2(g)$. The mixture is initially at equilibrium as described by the reaction

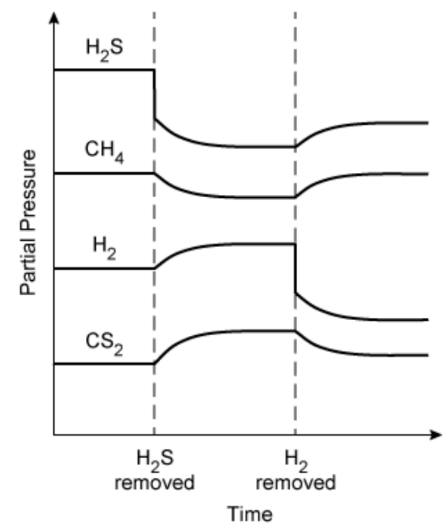
$$2\ H_2S(g) + CH_4(g) \rightleftarrows 4\ H_2(g) + CS_2(g)$$

A researcher then perturbs the equilibrium by removing some $H_2S(g)$ from the mixture. The mixture is allowed to reestablish equilibrium, and then the researcher removes some $H_2(g)$. Which graph best represents the partial pressures of each gaseous species as a function of time? (Note: Assume the volume stays constant.)

A.



B.



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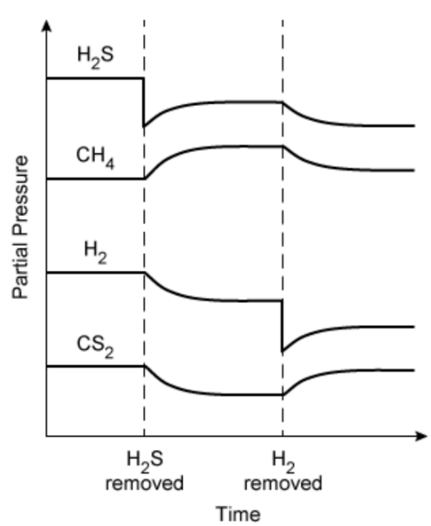
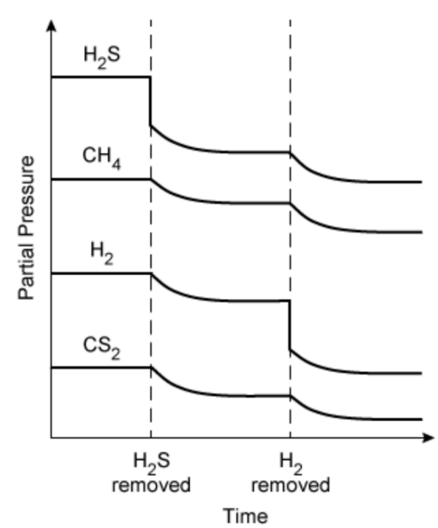


D.



3)

During deep underwater dives, scuba divers must overcome the external ambient pressure in order to expand their chests to inhale. To assist underwater breathing, air from scuba tanks is administered at high pressures to keep divers' lungs inflated. Scuba tanks are often filled with a mixture of N_2 and O_2 gas. Due to the increased pressure of the gases in the lungs, a greater amount of the gases are dissolved into the bloodstream during underwater dives. At a body temperature of 37°C, the solubility of N_2 gas in blood is 6.0×10^{-3} mol/L·atm and the solubility of O_2 gas is 1.3×10^{-2} mol/L·atm.

If a diver ascends too quickly after an extended dive, there is an increased risk of vascular bubble formation. Due to the decrease in ambient pressure, the excess gas dissolved in the blood may come out of solution as gas bubbles and cause a number of adverse effects. This condition is known as decompression sickness, or "the bends." To avoid this problem, scuba divers limit their rate of ascension to allow the excess gas dissolved in their blood to be safely removed through respiration. Because gas bubbles are good reflectors of sound, these bubbles can be detected using Doppler ultrasonic flowmetry after the diver has surfaced.

A diver is using a tank filled with 20% O_2 and 80% N_2 (by volume). If the pressure in the lungs is 5 atm, what is the expected equilibrium concentration of N_2 dissolved in the diver's blood? (Note: Assume that the mixtures of the gases in the lungs and tank are equal.)

A. 1.3×10^{-2} mol/L

B. 2.4×10^{-2} mol/L

C. 3.0×10^{-2} mol/L

D. 6.5×10^{-2} mol/L

4)

Consider the dimerization of nitrogen dioxide at room temperature:

$$2\ NO_2(g) \rightleftarrows N_2O_4(g) \quad K_p = 8\ \text{atm}^{-1}$$

If the equilibrium partial pressure of nitrogen dioxide in a container is 0.5 atm, what is the partial pressure of dinitrogen tetroxide?

A. 0.03 atm

B. 0.5 atm

C. 2 atm

D. 4 atm

5)

Based on the acid comparison in Table 1, which of the following acids will give a 0.10 M aqueous solution with the lowest pH?

Table 1 Comparison of Strengths of Selected 0.10 M Acids at 25°C

Acid	Acid Formula (HA)	Conjugate Base (A^-)	K_a
Hydrochloric	HCl	Cl^-	1.3×10^6
Carbonic	H_2CO_3	HCO_3^-	4.3×10^{-7}
Formic	HCO_2H	CO_2H^-	1.8×10^{-4}
Acetic	$HC_2H_3O_2$	$C_2H_3O_2^-$	1.8×10^{-5}
Lactic	$H_2C_3H_5O_3$	$HC_3H_5O_3^-$	1.4×10^{-4}
Ascorbic	$H_2C_6H_6O_6$	$HC_6H_6O_6^-$	7.9×10^{-5}
Citric	$H_3C_6H_5O_7$	$H_2C_6H_5O_7^-$	8.4×10^{-4}

A. Carbonic acid

B. Acetic acid

C. Lactic acid

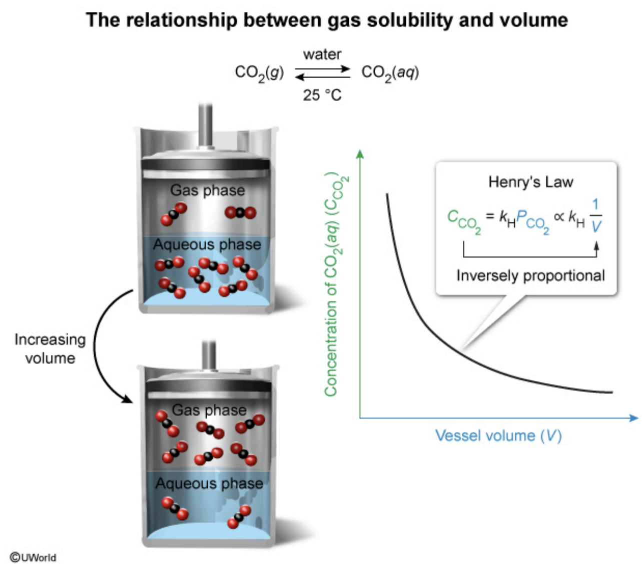
D. Citric acid

Answer Key

1) D
2) C
3) B
4) C
5) D

Explanations

1)



At a fixed temperature, the amount of gas dissolved in a particular solvent is directly proportional to the equilibrium partial pressure of that gas present above the solvent, as described by Henry's law of solubility:

$$C = k_{\text{H}}P_{\text{gas}}$$

where C is the concentration (or solubility) of the dissolved gas, k_H is the solubility coefficient specific to the gas and solvent at a given temperature, and P_{gas} is the partial pressure of the gas at equilibrium. If the pressure exerted by the gas on the solvent **increases** (or decreases), the amount of gas dissolved in the solvent will also **increase** (or decrease) by the same factor.

The question states that the volume of the reaction vessel containing an equilibrium mixture of dissolved $CO_2(aq)$ and $CO_2(g)$ is increased. According to Boyle's law, volume and pressure are inversely proportional:

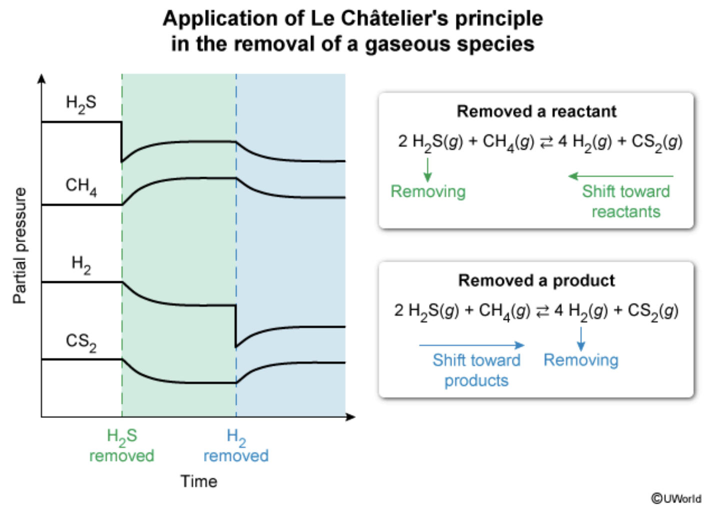
$$P \propto \frac{1}{V}$$

If volume **increases** (or decreases), pressure **decreases** (or increases) by the same factor. Substituting this relationship into Henry's law yields:

$$C \propto k_{\text{H}}\frac{1}{V}$$

Therefore, as the volume of the reaction vessel increases (ie, the partial pressure of $CO_2(g)$ decreases), the concentration of dissolved $CO_2(aq)$ decreases (ie, inversely proportional to volume), as depicted by the graph in Choice D.

2)



Le Châtelier's principle says that if the equilibrium state of a system is perturbed (eg, removal of a reactant or a product), the system will shift toward the reactants or toward the products to counteract the disturbance and establish a new equilibrium. **Removing products** causes the reaction to **form more products** (ie, the equilibrium shifts to the right), whereas **removing reactants** causes the reaction to **form more reactants** (ie, the equilibrium shifts to the left).

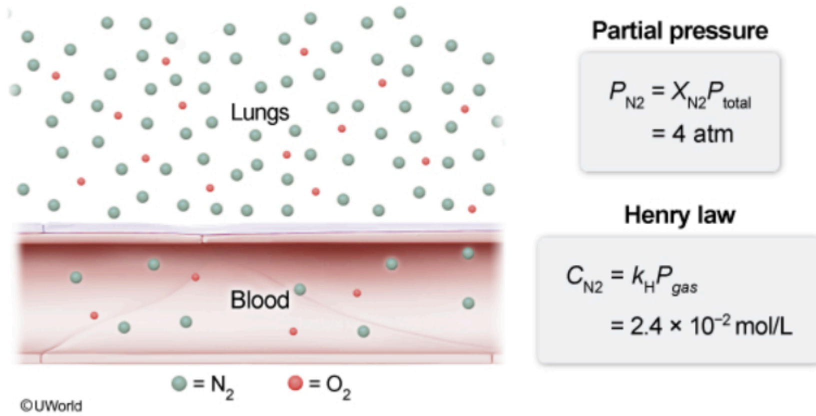
The question states that a mixture of the gases involved in the reaction

$$2\ H_2S(g) + CH_4(g) \rightleftarrows 4\ H_2(g) + CS_2(g)$$

is initially at equilibrium. A researcher then perturbs the equilibrium by first removing $H_2S(g)$ from the mixture; consequently, the partial pressure of $H_2S(g)$ decreases sharply. This causes the reverse reaction to be favored (ie, the reaction shifts toward reactants). In response, $H_2(g)$ and $CS_2(g)$ react to form $H_2S(g)$ and $CH_4(g)$, causing gradual increases in $H_2S(g)$ and $CH_4(g)$ partial pressures along with gradual decreases in $H_2(g)$ and $CS_2(g)$ partial pressures.

After allowing the mixture to reestablish equilibrium (ie, the partial pressures become constant as a function of time), some $H_2(g)$ is removed from the mixture. As such, a sharp decrease in the partial pressure of $H_2(g)$ occurs. This causes the forward reaction to be favored (ie, the reaction shifts toward products). In response, $H_2S(g)$ and $CH_4(g)$ react and gradually decrease in partial pressure, causing the partial pressures of $CS_2(g)$ and $H_2(g)$ to gradually increase. Therefore, the graph shown in Choice C best represents the partial pressures of the gaseous species after each perturbation.

3)



The amount of N_2 dissolved in the blood is in equilibrium with the partial pressure of N_2 in the gas phase. The **partial pressure** (P_{gas}) of a gas is the hypothetical pressure it would exert if it was the only gas occupying the same volume (Dalton law of partial pressures):

$$P_{\text{gas}} = X_{\text{gas}} P_{\text{total}}$$

where X_{gas} is the **mole fraction** of the gas, the ratio of the number of moles of a substance to the total number of moles present. An ideal gas's mole fraction is equal to its **fraction of the total volume**. Therefore, the partial pressure of N_2 is 80% of the total pressure (5 atm):

$$P_{\text{N2}} = (0.8)\,(5\text{ atm})$$

$$P_{\text{N2}} = 4\text{ atm}$$

The equilibrium between the gas partial pressure and the amount dissolved in solution is described by the Henry law of solubility:

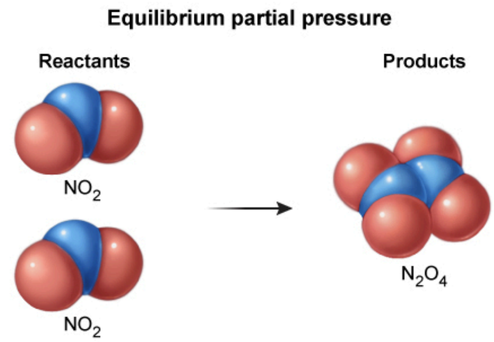
$$C = k_{\text{H}} P_{\text{gas}}$$

where C is the concentration of the dissolved gas, and k_H is the solubility coefficient specific to the gas and solution at a given temperature. Using 6.0 × 10^{-3} mol/atm for k_H and 4 atm for P, the concentration of dissolved N_2 is:

$$C = \left(6.0 \times 10^{-3}\,\frac{\text{mol}}{\text{atm} \cdot \text{L}}\right)(4\text{ atm})$$

$$C = 2.4 \times 10^{-2}\text{ mol/L}$$

4)



$$K_p = \frac{(P_{\text{products}})^m}{(P_{\text{reactants}})^n} = \frac{P_{N_2O_4}}{P_{NO_2} \times P_{NO_2}} = \frac{P_{N_2O_4}}{(P_{NO_2})^2}$$

The **equilibrium constant** for any chemical reaction is the **molar ratio of products to reactants** when the system is equilibrium. The partial pressure of a gas is directly proportional to the number of moles of that gas according to the ideal gas law; therefore, partial pressure can be substituted for molar concentration. If more than one of a given molecule is involved in a reaction, each molecule involved is counted separately. Therefore, each product or reactant must be **raised to the power** of its stoichiometric coefficient to account for all molecules involved in the reaction.

Nitrogen dioxide (NO_2) can combine with itself to form dinitrogen tetroxide (N_2O_4). The equilibrium constant for this reaction (K_p = 8 atm^{-1}) is the ratio of the partial pressure of N_2O_4 to that of NO_2 when equilibrium is achieved. However, because two NO_2 molecules are required for each reaction, the NO_2 partial pressure (P_{NO_2} = 0.5 atm) must be squared. Therefore, the equilibrium partial pressures of NO_2 and N_2O_4 in the described scenario are related by the equation:

$$K_p = \frac{P_{N_2O_4}}{\left(P_{NO_2}\right)^2}$$

Substituting the values into the equation yields:

$$8 \text{ atm}^{-1} = \frac{P_{N_2O_4}}{(0.5 \text{ atm})^2} = \frac{P_{N_2O_4}}{0.25 \text{ atm}^2}$$

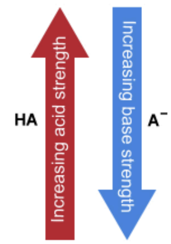
This equation can be rearranged and solved as follows:

$$P_{N_2O_4} = 0.25 \text{ atm}^2 \times 8 \text{ atm}^{-1} = 2 \text{ atm}$$

5)

$HA \leftrightarrows H^+ + A^-$ $\quad K_a = \frac{[H^+][A^-]}{[HA]}$ $\quad pK_a = -\log(K_a)$ $\quad pH = -\frac{1}{2}\log([HA]K_a)$

Acid	HA	A⁻	K_a	pK_a	pH
Citric	$H_3C_6H_5O_7$	$H_2C_6H_5O_7^-$	8.4×10^{-4}	3.1	2.04
Formic	HCO_2H	CO_2H^-	1.8×10^{-4}	3.7	2.37
Lactic	$H_2C_3H_5O_3$	$HC_3H_5O_3^-$	1.4×10^{-4}	3.9	2.43
Ascorbic	$H_2C_6H_6O_6$	$HC_6H_6O_6^-$	7.9×10^{-5}	4.1	2.55
Acetic	$HC_2H_3O_2$	$C_2H_3O_2^-$	1.8×10^{-5}	4.7	2.87
Carbonic	H_2CO_3	HCO_3^-	4.3×10^{-7}	6.4	3.68



When an acid (HA) is dissolved and ionized in water, a proton (H^+ ion) and a conjugate base of the acid (A^-) are released into solution.

$$HA\ (aq) \underset{}{\overset{K_a}{\rightleftarrows}} H^+\ (aq) + A^-\ (aq)$$

For this process, strong acids have an equilibrium favoring the products (ie, the formation of H^+ and A^-) whereas weak acids have an equilibrium favoring the reactant (non-ionized acid, HA). Consequently, stronger acids produce higher concentrations of H^+ and A^- in solution relative to non-ionized HA. This equilibrium is quantified by the **acid ionization constant K_a**, which is larger for stronger acids.

$$K_a = \frac{[H^+]\,[A^-]}{[HA]}$$

Due to the wide range in K_a values, the **pK_a** scale is frequently used as a convenient way of comparing acid strength.

$$pK_a = -\log K_a$$

Because of the logarithmic definition of the pK_a scale, **acid strength increases as pK_a *decreases***. Analogously, the pH of a solution is defined as:

$$pH = -\log\,[H^+]$$

Accordingly, the acid that most favors H^+ production yields the lowest pH. Therefore, when compared in equal amounts, the acid with the **smallest pK_a (largest K_a)** is the **strongest acid** and yields the solution with the **highest [H⁺]** and the **lowest pH**.

Of the acids compared in Table 1, citric acid is the most acidic (ie, largest K_a, smallest pK_a) and will produce a solution with the lowest (most acidic) pH.