1. For the reaction H_2 (g) + $Br_2 \rightarrow 2$ HBr (g), if the

initial rate of formation of hydrogen bromide is 4.2×10^{-3} mol dm^3 s^{-1} what is the initial rate of

consumption of H_2?

- ○ 4.2×10^{-3} mol dm^{-3} s^{-1}
- ◉ 2.1×10^{-3} mol dm^{-3} s^{-1}
- ○ 8.4×10^{-3} mol dm^{-3} s^{-1}
- ○ Not possible to determine.

✓ **Correct**

Use the stoichiometry of the reaction to determine the relative rates of formation and consumption. As 2 molecules of HBr are formed for every one of H_2 consumed the rate of consumption will be half.

2. For a general reaction $3A + 2B \rightarrow$ products, the rate equation takes the form: rate of reaction $= k[A]^1[B]^2$. Which of the following are

correct?

- [ ] The overall order of reaction $= 2$

- [x] The order of reaction with respect to $A = 1$

Correct

The overall order of reaction $= (2 + 1) = 3$, the units of

the rate constant $k = \mathrm{dm^6\ mol^{-2}\ s^{-1}}$.

- [x] The order of the reaction with respect to $B = 2$

Correct

The overall order of reaction $= (2 + 1) = 3$, the units of

the rate constant $k = \mathrm{dm^6\ mol^{-2}\ s^{-1}}$.

- [ ] The units of the rate constant $k = \mathrm{s^{-1}}$

3. For the general reaction $2\mathrm{A} + \mathrm{B} \rightarrow$ products, the

rate equation takes the following form: $\text{rate of reaction} = k[\mathrm{A}]^2[\mathrm{B}]^1$. What are the units of the rate constant, k?

- ○ s^{-1}
- ○ $\mathrm{dm}^3\ \mathrm{mol}^{-1}\ \mathrm{s}^{-1}$
- ○ $\mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}$
- ◉ $\mathrm{dm}^6\ \mathrm{mol}^{-2}\ \mathrm{s}^{-1}$

Correct

In order to find the rate constant rearrange the rate of reaction

equation such that the rate constant, k, is the subject of the

equation:

$$k = \frac{\text{rate of reaction}}{[\mathrm{A}]^2[\mathrm{B}]}$$

Thus, the

units of k are given by:

$$\frac{\mathrm{mol\ dm}^{-3}\ \mathrm{s}^{-1}}{(\mathrm{mol\ dm}^{-3})(\mathrm{mol\ dm}^{-3})(\mathrm{mol\ dm}^{-3})} = \frac{\mathrm{s}^{-1}}{(\mathrm{mol\ dm}^{-3})(\mathrm{mol\ dm}^{-3})} = \mathrm{dm}^6\ \mathrm{mol}^{-2}\ \mathrm{s}^{-1}$$

4. The rate of consumption of reactants and products in the following

reaction: $2\,\mathrm{I}\ (\mathrm{g}) + \mathrm{Ar}\ (\mathrm{g}) \rightarrow \mathrm{I}_2(\mathrm{g}) + \mathrm{Ar}\ (\mathrm{g})$ can be defined

by which of the following expressions:

- (selected) rate of reaction: $-\frac{1}{2}\frac{\mathrm{d[I]}}{\mathrm{dt}} = -\frac{\mathrm{d[Ar]}}{\mathrm{dt}} = \frac{\mathrm{d[I_2]}}{\mathrm{dt}} = \frac{\mathrm{d[Ar]}}{\mathrm{dI}}$
- rate of reaction: $\frac{1}{2}\frac{\mathrm{d[I]}}{\mathrm{dt}} = \frac{\mathrm{d[Ar]}}{\mathrm{dt}} = \frac{\mathrm{d[I_2]}}{\mathrm{dt}} = \frac{\mathrm{d[Ar]}}{\mathrm{dI}}$
- rate of reaction: $\frac{\mathrm{d[I]}^2}{\mathrm{dt}} = \frac{\mathrm{d[Ar]}}{\mathrm{dt}} = \frac{\mathrm{d[I_2]}}{\mathrm{dt}} = \frac{\mathrm{d[Ar]}}{\mathrm{dI}}$
- rate of reaction: $\frac{2\mathrm{d[I]}}{\mathrm{dt}} = \frac{\mathrm{d[Ar]}}{\mathrm{dt}} = \frac{\mathrm{d[I_2]}}{\mathrm{dt}} = \frac{\mathrm{d[Ar]}}{\mathrm{dI}}$

✓ **Correct**

The rate of reaction can be defined such that no matter whether the

sonsuption of a reactiont or formation of a product is followed

experimentally, the same value of rate is given. For a general

reaction:

$a\mathrm{A} + b\mathrm{B} \rightarrow p\mathrm{P} + q\mathrm{Q}$ rate of reaction: $-\frac{1}{a}\frac{\mathrm{d[A]}}{\mathrm{dt}} = -\frac{1}{b}\frac{\mathrm{d[B]}}{\mathrm{dt}} = \frac{1}{p}\frac{\mathrm{d[P]}}{\mathrm{dt}} = \frac{1}{q}\frac{\mathrm{d[Q]}}{\mathrm{dt}}$

, the terms $\frac{1}{a}, \frac{1}{b}$

are added in order to take account of the

stoichiometry of the reaction. A minus sign is placed before the rate

of consumption of reactant expressions.

5. The expression for the rate of reaction for the following elementary

reaction $\mathrm{NO} + \mathrm{O}_3 \rightarrow \mathrm{NO}_2 + \mathrm{O}_2$ is:

- (selected) $\text{rate of reaction} = k[\mathrm{NO}][\mathrm{O}_3]$
- $\text{rate of reaction} = k[\mathrm{NO}][\mathrm{O}_3]^3$
- $\text{rate of reaction} = k[\mathrm{NO}_2][\mathrm{O}_2]$
- $\text{rate of reaction} = k\frac{[\mathrm{NO}_2][\mathrm{O}_2]}{[\mathrm{NO}][\mathrm{O}_3]}$

Correct

Since this is an elementary reaction, it is possible to use the

stoichiometric chemical equation to write the rate equation. As an

elementary process, the reactants will appear in the rate equation and

not the products. The rate equation, contains the order of each

reactant and in this case, this would be equal to the stoichiometric

coefficient thus $\text{rate of reaction} = k[\mathrm{NO}][\mathrm{O}_3]$

6. For the following gas-phase bimolecular elementary reaction

$2\mathrm{Cl}^\bullet \rightarrow \mathrm{Cl}_2$. Which rate

equation, expressed in terms of a differential, shows the relationship

between the rate of consumption of reactant and rate of reaction?

- ○ $-2\frac{\mathrm{d}[\mathrm{Cl}^\bullet]}{\mathrm{d}t} = k[\mathrm{Cl}^\bullet]$
- ○ $-2\frac{\mathrm{d}[\mathrm{Cl}^\bullet]}{\mathrm{d}t} = k[\mathrm{Cl}^\bullet]^2$
- ○ $-\frac{\mathrm{d}[\mathrm{Cl}^\bullet]}{\mathrm{d}t} = k[\mathrm{Cl}^\bullet]^2$
- ◉ $-\frac{\mathrm{d}[\mathrm{Cl}^\bullet]}{\mathrm{d}t} = 2 \times k[\mathrm{Cl}^\bullet]^2$

✓ **Correct**

Since this is an elementary reaction, and two Cl radicals are

involved in the reaction (as revealed by the stoichiometric

coefficient), this is given by $\text{rate of reaction} = -\frac{1}{2}\frac{\mathrm{d}[\mathrm{Cl}^\bullet]}{\mathrm{d}t}$. Furthermore,

$\text{rate of reaction} = -\frac{1}{2}\frac{\mathrm{d}[\mathrm{Cl}^\bullet]}{\mathrm{d}t} = \frac{\mathrm{d}[\mathrm{Cl}_2]}{\mathrm{d}t}$ and

$\frac{\mathrm{d}[\mathrm{Cl}_2]}{\mathrm{d}t} = k[\mathrm{Cl}^\bullet]^2$.

Hence, $-\frac{\mathrm{d}[\mathrm{Cl}^\bullet]}{\mathrm{d}t} = 2 \times k[\mathrm{Cl}^\bullet]^2$.

7. $2N_2O_5$ (g) $\rightarrow 4\ NO_2$ (g) $+ O_2$ (g) follows first order reaction

kinetics, where the $\text{rate of reaction} = k[N_2O_5]$. The integrated form of the rate equation is:

- ○ $[N_2O_5] = [N_2O_5]_0 - 2kt$
- ◉ $\ln[N_2O_5] = \ln[N_2O_5]_0 - 2kt$
- ○ $2[N_2O_5] = [N_2O_5]_0 - kt$
- ○ $2\ln[N_2O_5] = \ln[N_2O_5]_0 - kt$

✓ **Correct**

The integrated rate equation links the concentration of the reactant

N_2O_5 at a particular time $[N_2O_5]_t$, to the time, t, the concentration at the start of

the reaction $[N_2O_5]_0$ and the rate contant,

k. Since the rate of decomposition of $N_2O_5 = k[N_2O_5]$, and using the standard integral:

$$\int_{x_1}^{x_2} \frac{\mathrm{dx}}{\mathrm{x}} = [\ln \mathrm{x}]_{x_1}^{x_2} = \ln \mathrm{x}_2 - \ln \mathrm{x}_1$$

after integrating and rearranging, the following relation is found:

$$\ln[N_2O_5] = \ln[N_2O_5]_0 - 2kt$$

8. For the reaction $2\ N_2O_5\ (g) \rightarrow 4\ NO_2\ (g) + O_2\ (g)$, the rate

constant, $k = 3.38 \times 10^{-5}\ s^{-1}$. What is the

overall reaction order?

- ◯ Zero
- ◉ First
- ◯ Second
- ◯ Third

✓ **Correct**

Given that $k = 3.38 \times 10^{-5}\ s^{-1}$, with units of

s^{-1}, the overall oder of the reaction can be found since

the units of the rate constant changes with the overall order. Thus,

in this case, the reaction is first order.

If $\text{rate} = k[\text{A}]$

Rearranging:

$\frac{\text{rate}}{[\text{A}]} = k$

$\frac{\text{mol dm}^{-3}\ \text{s}^{-1}}{\text{mol dm}^{-3}} = k$

$k = \text{s}^{-1}$

A zero order reaction would have a rate contant, k, with units

$\text{mol dm}^{-3}\ \text{s}^{-1}$.

A second order reaction would have a rate contant, k, with units

$\text{dm}^{-3}\ \text{mol}^{-1}\ \text{s}^{-1}$.

A third order reaction would have a rate contant, k, with units

$\text{dm}^{-6}\ \text{mol}^{-2}\ \text{s}^{-1}$.

9. For a first order elementary reaction $\mathrm{A} \rightarrow \text{products}$, the rate of reaction $= k[\mathrm{A}]$. Please select all that apply.

- [ ] $t_{1/2} = \frac{1}{2k[\mathrm{A}]_0}$
- [x] $t_{1/2} = \frac{\ln 2}{k}$

✓ **Correct**

There are a number of quantities such as half-life ($t_{1/2}$) and the change in concentration with respect to time (integrated first order rate equation) that can be used to understand the kinetics of a chemical reaction and monitor how much reactant will be used up as the chemical reaction progresses. For first order reactions, the half-life ($t_{1/2}$) does not depend on initial concentration $[\mathrm{A}]_0$ of the reactant, rather it is constant through the course of the reaction. The integrated rate equation allows the reactant concentration at any point in time of the reaction progress to be monitored. Whilst a plot of $\ln[\mathrm{A}]$ against time will give a straight-line graph that confirms the order of the reaction.

- [x] $\ln[\mathrm{A}]_t = \ln[\mathrm{A}]_0 - k\mathrm{t}$

Correct

There are a number of quantities such as half-life ($t_{1/2}$) and the change in concentration with respect to time (integrated first order rate equation) that can be used to understand the kinetics of a chemical reaction and monitor how much reactant will be used up as the chemical reaction progresses. For first order reactions, the half-life ($t_{1/2}$) does not depend on initial concentration $[\mathrm{A}]_0$ of the reactant, rather it is constant through the course of the reaction. The integrated rate equation allows the reactant concentration at any point in time of the reaction progress to be monitored. Whilst a plot of $\ln[\mathrm{A}]$ against time will give a straight-line graph that confirms the order of the reaction.

- [ ] A plot of $[\mathrm{A}]$ versus time, t, will give a straight-line graph.

10. The first order gas-phase process: $C_2H_6 \rightarrow 2\ CH_3$ has a rate constant, $k = 5.36 \times 10^{-4}\ s^{-1}$ at $700°C$. If in a reaction vessel, the initial

concentration of ethane was $0.05\ \mathrm{mol\ dm}^{-3}$, what is its

concentration after $250\ \mathrm{s}$?

- ○ -0.084
- ○ 0.034
- ◉ 0.044
- ○ -3.13

✓ **Correct**

The integrated rate equation for a first order reaction is:

$\ln[\mathrm{A}]_t = \ln[\mathrm{A}]_0 - kt$. It is possible to

find the concentration of the reactant, ethane, at any time during the

reaction, using the alternative way of writing integrated rate

equation for a first oder reaction: $[\mathrm{A}]_t = [\mathrm{A}]_0\ \mathrm{e}^{-kt}$.