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# Collision theory worksheet 16. 1 answer key free printable version 2

By the end of this section, you will be able to: Use the postulates of collision theory to explain the effects of physical state, temperature, and concentration on reaction rates Define the concepts of activation energy and transition state Use the Arrhenius equation in calculations relating rate constants to temperature We should not be surprised that atoms, molecules, or ions must collide before they can react with each other. Atoms must be close together to form chemical bonds. This simple premise is the basis for a very powerful theory that explains many observations regarding chemical kinetics, including factors affecting reaction rates. Collision theory is based on the following postulates: The rate of a reaction is proportional to the rate of reactant collisions:  $[text{rate}][text{reaction}:rate] \cdot [text{propo}][text{collisions}] \cdot [text{time}][text{rate}]$ . The reacting species must collide in an orientation that allows contact between the atoms that will become bonded together in the product. The collision must occur with adequate energy to permit mutual penetration of the reacting species. AAvalence shells so that the electrons can rearrange and form new bonds (and new chemical species). We can see the importance of the two physical factors noted in postulates 2 and 3, the orientation and energy of collisions, when we consider the reaction of carbon monoxide with oxygen:  $[text{rate}][text{CO}(g) + text{O}_2(g) \rightarrow text{CO}_2(g)] \cdot [text{longrightarrow CO}_2 \cdot [text{CO}_2(g)]$ . Carbon monoxide is a pollutant produced by the combustion of hydrocarbon fuels. To reduce that pollutant, automobiles have catalytic converters that use a catalyst to carry out this reaction. It is a side reaction of the combustion of gunpowder that results in muzzle flash for many firearms. If carbon monoxide and oxygen are present in sufficient quantity, the reaction is spontaneous at high temperature and pressure. The first step in the gas-phase reaction carbon and oxygen monoxide is a collision between the two mollets:  $[text{rate}][text{CO} + text{O}_2 \rightarrow text{CO}_2 + text{O}]$ . In the first case, the oxygen side of the monomer of carbon monoxide collides with the oxygen atom. In the second case, the carbon side of the non-carbon monomeric is collides with the oxygen atom. The second case is clearly more likely to result in the form of carbon dioxide, which has a central carbon attached to two oxygen atoms. This is a quite simple example of how the collision guidance is important in terms of creating the desired product. Figure 1. Illustrated are two collisions that can proceed in the front direction, producing an appropriate activated complex along the way. As Figure 2 demonstrates, even a collision with the correct orientation can no longer form the product of reaction. In the study of the mechanisms of reactions, each of these transitions of atoms is called a state of a transaction or proposed activated complex.

Figure 2. Possible states of transactions (activated complexes) for carbon monoxide reacting with oxygen. The first diagram shows the first two examples of successful collisions that will lead to the formation of carbon dioxide. The second diagram shows the failure of a collision due to incorrect orientation. The third diagram shows the failure of a collision due to insufficient energy.

The first diagram shows the reaction of carbon monoxide with oxygen. The reactants are carbon monoxide (CO) and oxygen (O<sub>2</sub>). The products are carbon dioxide (CO<sub>2</sub>) and oxygen (O). The reaction is shown as CO + O<sub>2</sub> → CO<sub>2</sub> + O. The first collision is successful because the CO molecule is oriented correctly relative to the O<sub>2</sub> molecule. The second collision is unsuccessful because the CO molecule is oriented incorrectly relative to the O<sub>2</sub> molecule.

The second diagram shows the reaction of carbon monoxide with oxygen. The reactants are carbon monoxide (CO) and oxygen (O<sub>2</sub>). The products are carbon dioxide (CO<sub>2</sub>) and oxygen (O). The reaction is shown as CO + O<sub>2</sub> → CO<sub>2</sub> + O. The collision is unsuccessful because the CO molecule is oriented incorrectly relative to the O<sub>2</sub> molecule.

The third diagram shows the reaction of carbon monoxide with oxygen. The reactants are carbon monoxide (CO) and oxygen (O<sub>2</sub>). The products are carbon dioxide (CO<sub>2</sub>) and oxygen (O). The reaction is shown as CO + O<sub>2</sub> → CO<sub>2</sub> + O. The collision is unsuccessful because the CO molecule does not have enough energy to overcome the potential energy barrier.

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