ENERGY CHANGES & **RATES OF REACTION**

STUDYING THE RATES OF CHEMICAL REACTIONS IS KNOWN AS "KINETICS"

WHAT IS KINETICS?

- THE RATE AT WHICH CHEMICAL REACTIONS OCCUR
  - THE CHANGE IN CONCENTRATION OF REACTANTS OVER TIME
  - THE CHANGE IN CONCENTRATION OF PRODUCTS OVER TIME

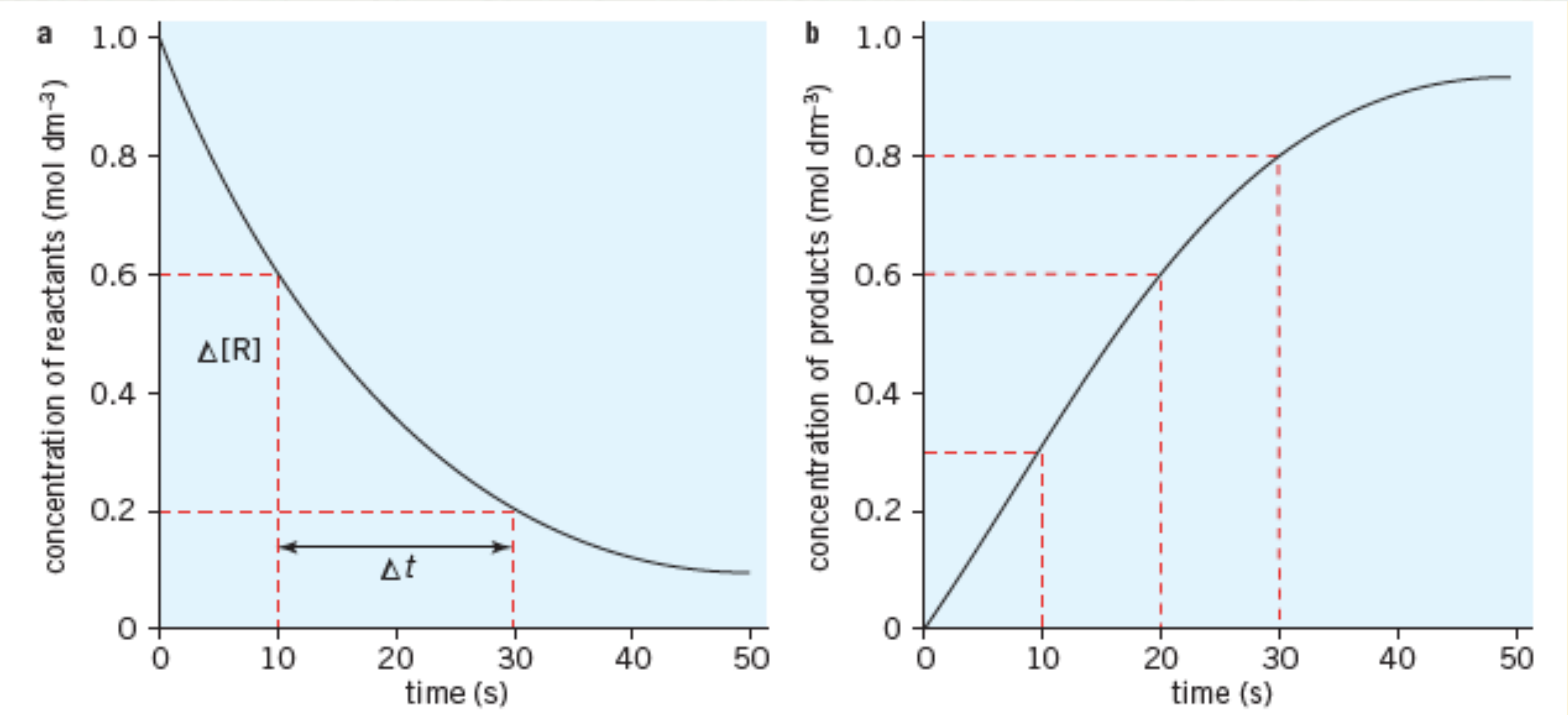


Figure 7.1.1 As a reaction proceeds (a) the concentration of reactants decreases and (b) the concentration of products increases.

HOW DO WE MEASURE RATES?

Chemical reactions indicate the overall change that is observed. Most reactions take place through a series of steps which are usually too quick to observe.

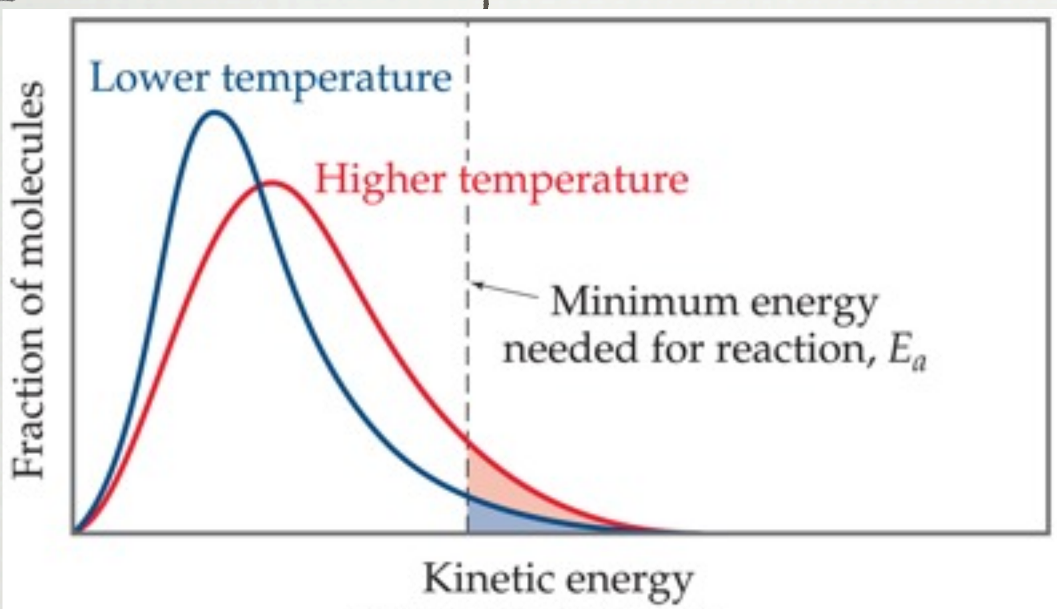
- ☐ CHANGE IN MASS
- ☐ CHANGE IN CONCENTRATION
- ☐ CHANGE IN VOLUME
- ☐ CHANGE IN PRESSURE
- ☐ CHANGE IN COLOUR
- ☐ CHANGE IN CONDUCTIVITY
- ☐ CHANGE IN LIGHT ABSORPTION

what affects rates?

- Temperature
- Concentration of Reactants
- Surface area
- Catalysts
- The Nature of the Reactants
  - Chemical compounds vary considerably in their chemical reactivities

but why?

- Increase temperature, increase rate of reaction & vice versa



- At higher temperature, molecules have more energy
- therefore, more molecules will have enough energy to overcome Ea and to form products

Temperature & Rates

effect of concentration

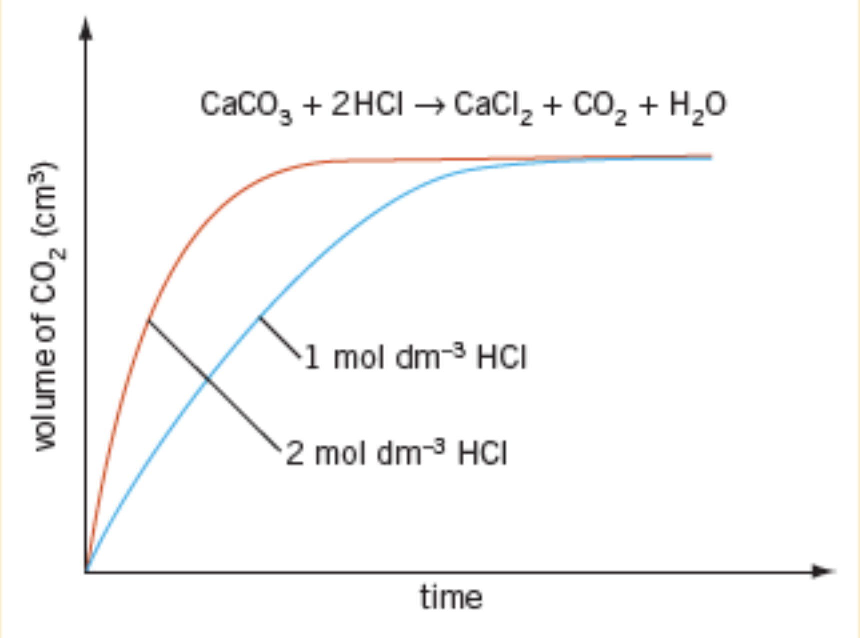


Figure 7.2.10 Increasing the concentration of an aqueous reactant increases the average rate of the reaction.

- Increase concentration, increase rate of reaction & vice versa
- Recall - concentration is mol/volume (c = n/V)
- increasing pressure of a gas has the same effect as increasing concentration:
  - more particles in a particular space means more chances of colliding

effect of surface area

- increase surface area, increase rate of reaction & vice versa
- what dissolves faster: a lump of sugar or a spoonful of fine sugar?
- more surface area will give more opportunities for the reaction to take place. What about particle size???

effect of catalysts

- a catalyst is a substance that speeds up a reaction without actually being used up itself
- catalysts provide an alternate pathway for the reaction with lower energy

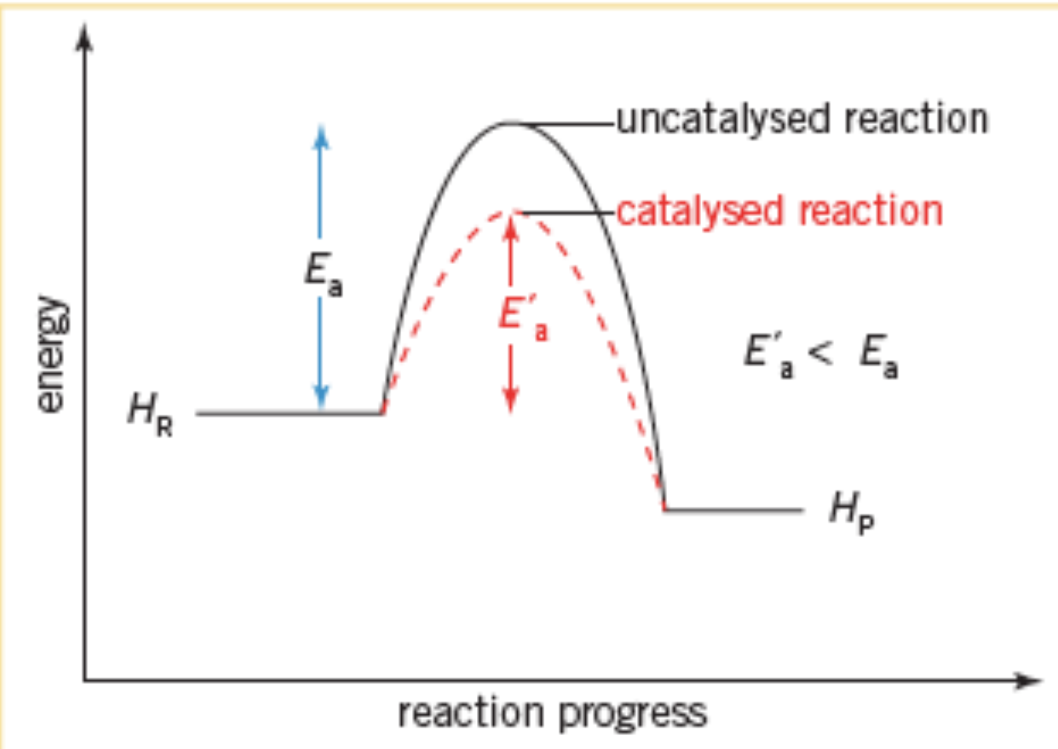


Figure 7.2.16 Catalysts lower the activation energy (E_a) and so increase the rate of reaction.

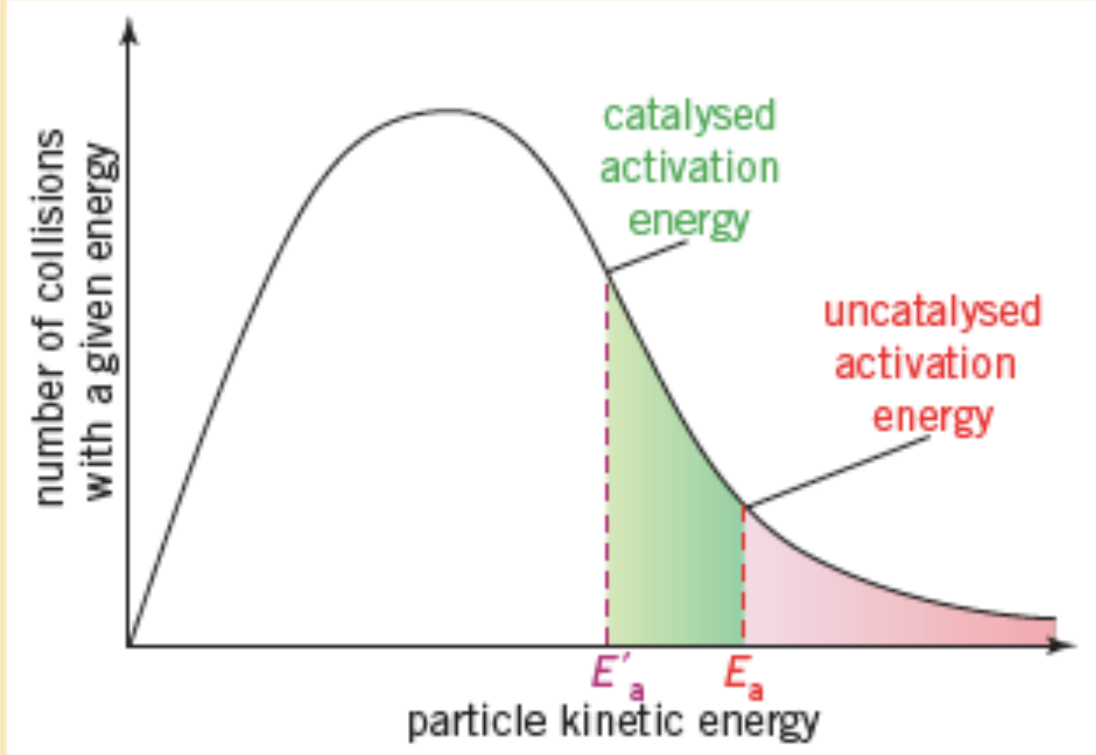
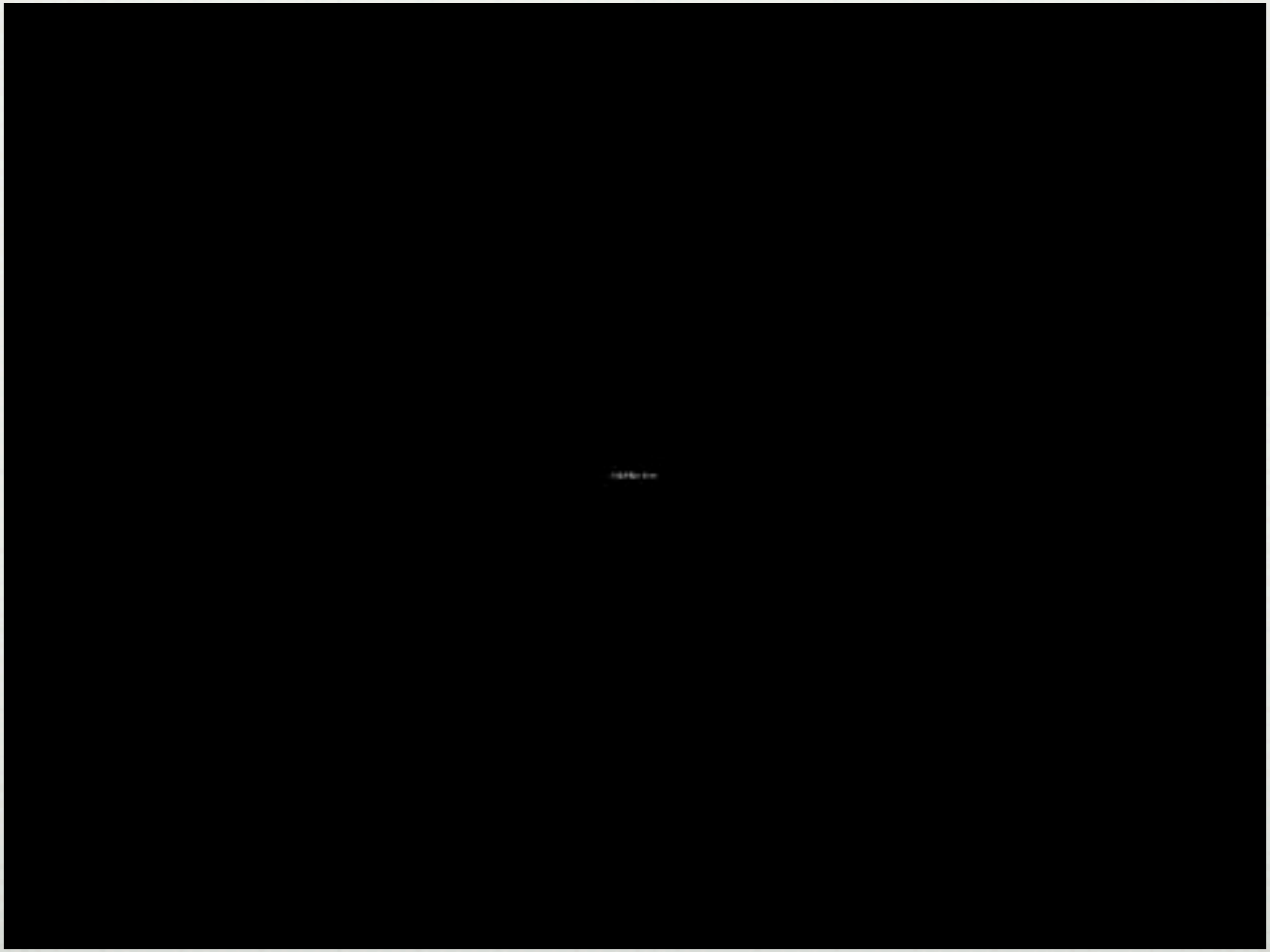
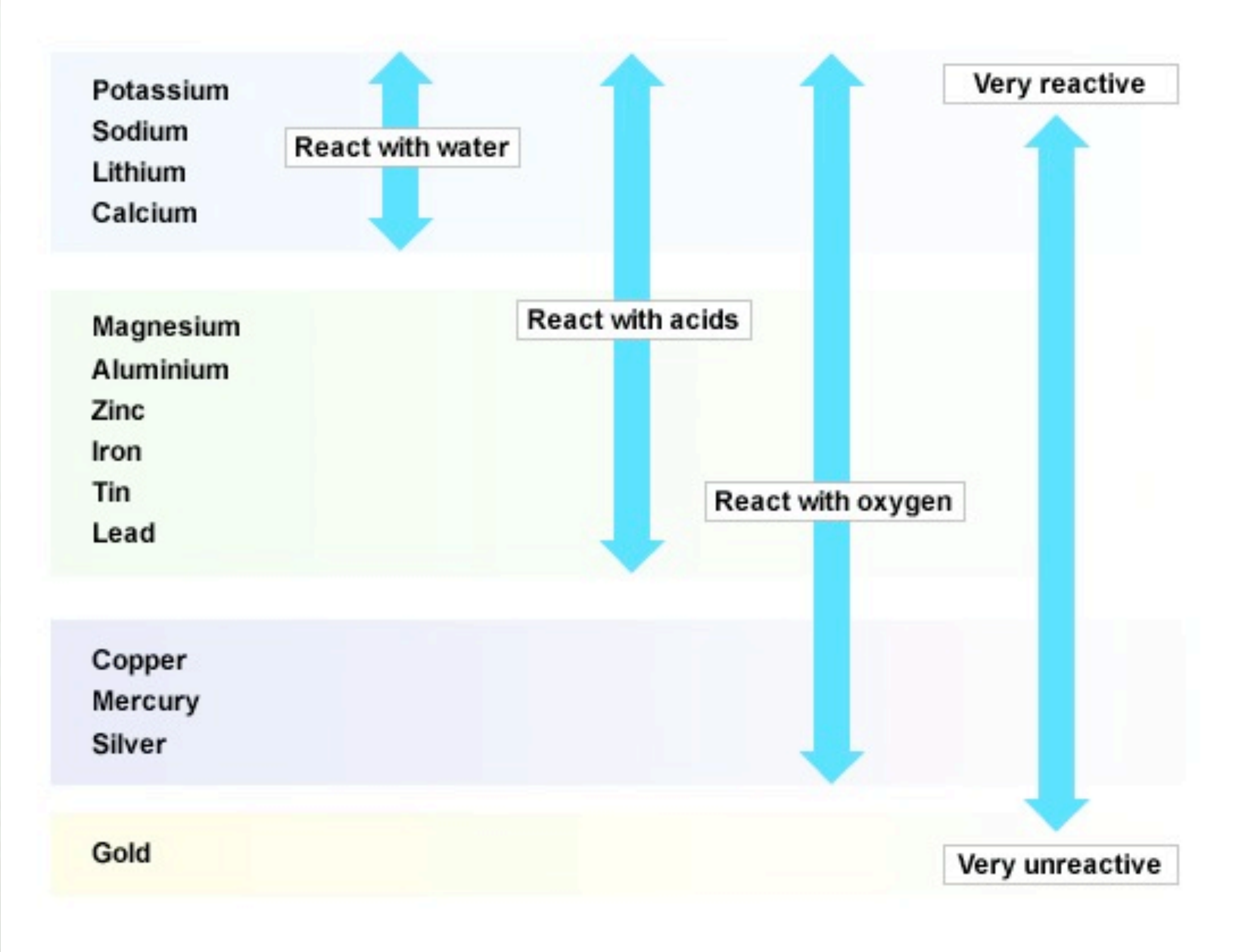


Figure 7.2.17 Catalysts lower the activation energy (E_a) so that more collisions will result in product formation.

What catalysts can do!



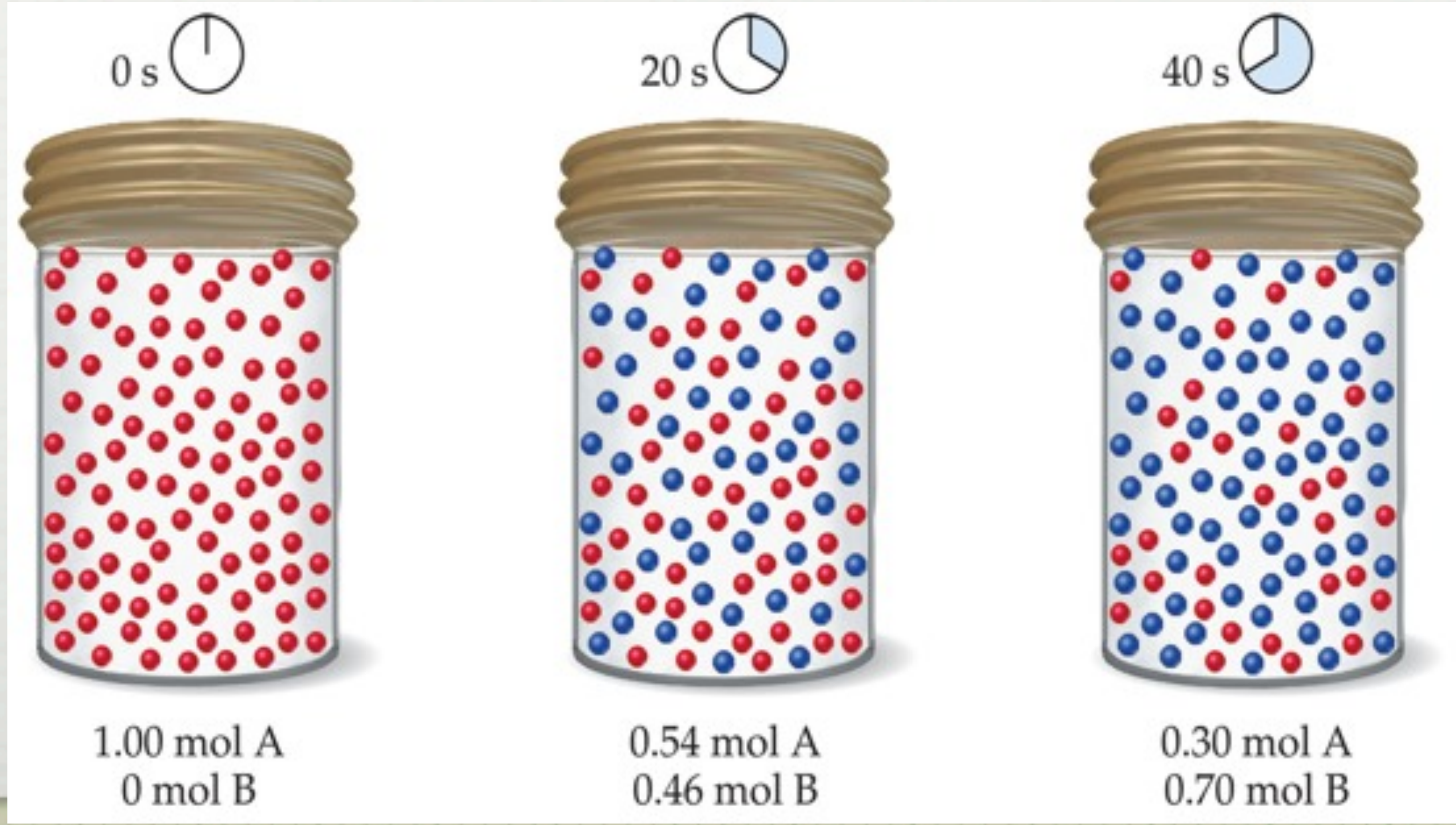
Nature of Reactants



REACTION RATES

- THE RATE OF A CHEMICAL REACTION CAN BE DETERMINED BY MONITORING THE CHANGE IN CONCENTRATION OF EITHER REACTANTS DISAPPEARING OR BY THE PRODUCTS APPEARING AS A FUNCTION OF TIME.

- REACTION RATE = $-[A] / \Delta T$ OR $[B] / \Delta T$



Example 1: Reaction Rates

$C_4H_9Cl(aq) + H_2O(l) \longrightarrow C_4H_9OH(aq) + HCl(aq)$

Time, t(s)	$[C_4H_9Cl]$ M
0.0	0.1000
50.0	0.0905
100.0	0.0820
150.0	0.0741
200.0	0.0671
300.0	0.0549
400.0	0.0448
500.0	0.0368
800.0	0.0200
10,000	0

In this reaction, the concentration of butyl chloride, C_4H_9Cl, was measured at various times, t.

$$\text{Rate} = \frac{-\Delta[C_4H_9Cl]}{\Delta t}$$

- Note: by convention, rates are positive. So, if you are working with reactants disappearing, you must multiply by -1!

Calculating Reaction Rates

$C_4H_9Cl(aq) + H_2O(l) \longrightarrow C_4H_9OH(aq) + HCl\ (aq)$

Time, t(s)	$[C_4H_9Cl]$ (M)	Average Rate, M/s
0.0	0.1000	
		1.9×10^{-4}
50.0	0.0905	
		1.7×10^{-4}
100.0	0.0820	
		1.6×10^{-4}
150.0	0.0741	
		1.4×10^{-4}
200.0	0.0671	
		1.22×10^{-4}
300.0	0.0549	
		1.01×10^{-4}
400.0	0.0448	
		0.80×10^{-4}
500.0	0.0368	
		0.560×10^{-4}
800.0	0.0200	
10,000	0	

The **average rate** of the reaction over each interval is the change in concentration divided by the change in time

The most common method of changing a reaction rate is through changing the concentration of reactants.

Mathematically:

$$\text{rate} = \frac{\Delta\text{concentration}}{\Delta\text{time}}$$

Units?

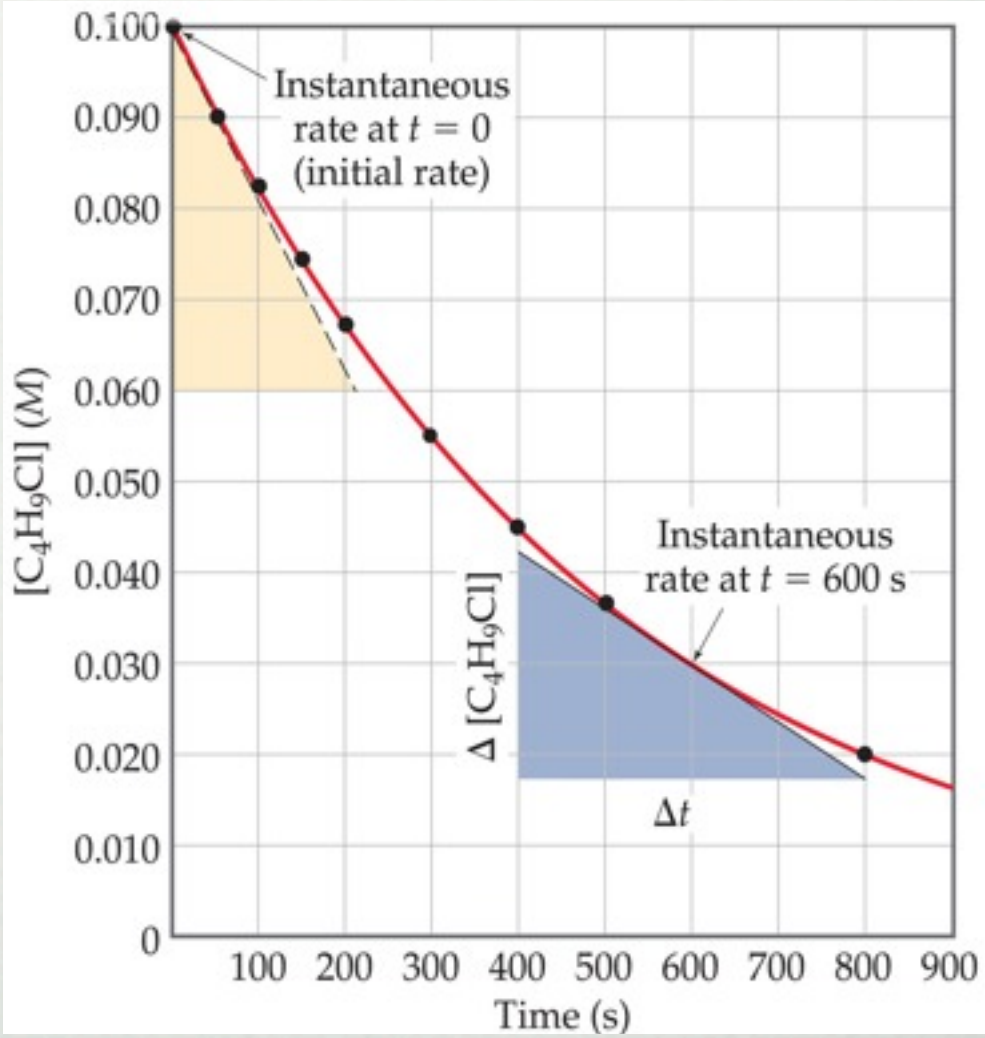
mol / s

mol / L • s = M / s

Reaction Rates

$C_4H_9Cl(aq) + H_2O(l) \longrightarrow C_4H_9OH(aq) + HCl(aq)$

- A plot of concentration vs. time for this reaction yields a curve like this.
- The slope of a line tangent to the curve at any point is the instantaneous rate at that time.
- Secants yield the average rate



Reaction Rates

$C_4H_9Cl(aq) + H_2O(l) \longrightarrow C_4H_9OH(aq) + HCl(aq)$

- The reaction slows down with time because the concentration of the reactants decreases.
- So there are less molecules to collide and react

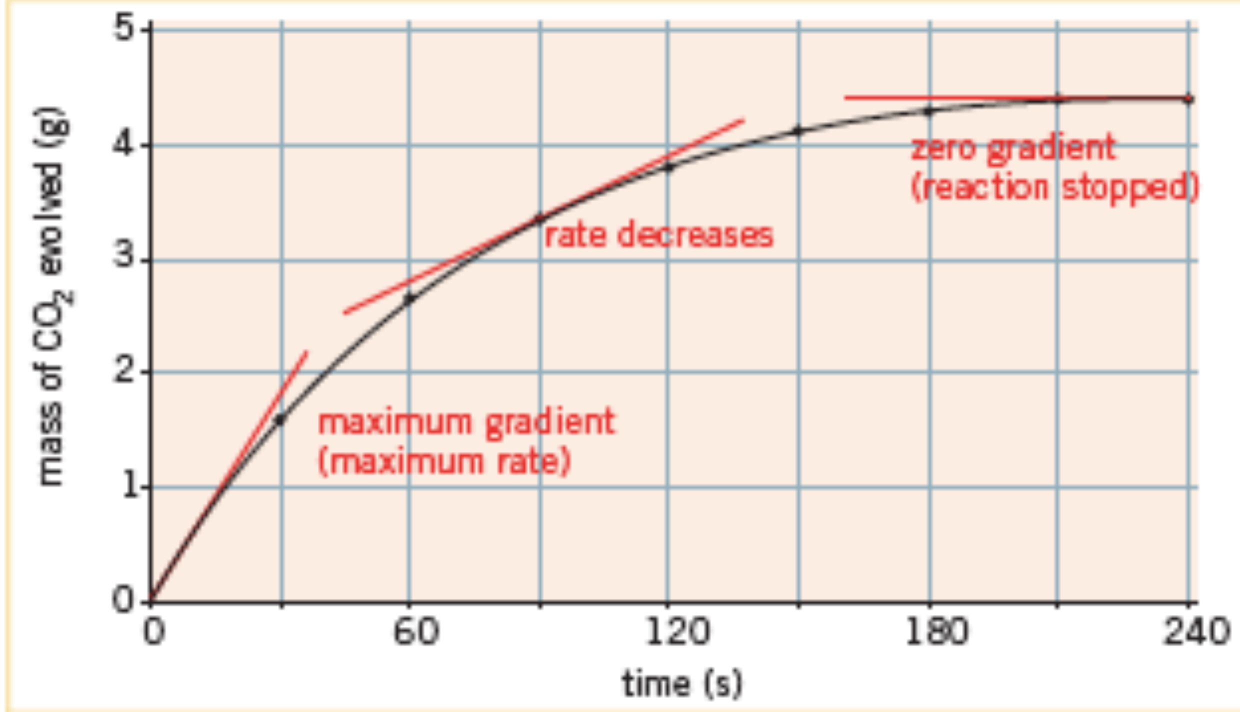


Figure 7.1.4 Graph of reaction of $CaCO_3$ with excess HCl to produce CO_2.

Reaction Rate Determination

$C_4H_9Cl(aq) + H_2O(l) \longrightarrow C_4H_9OH(aq) + HCl(aq)$

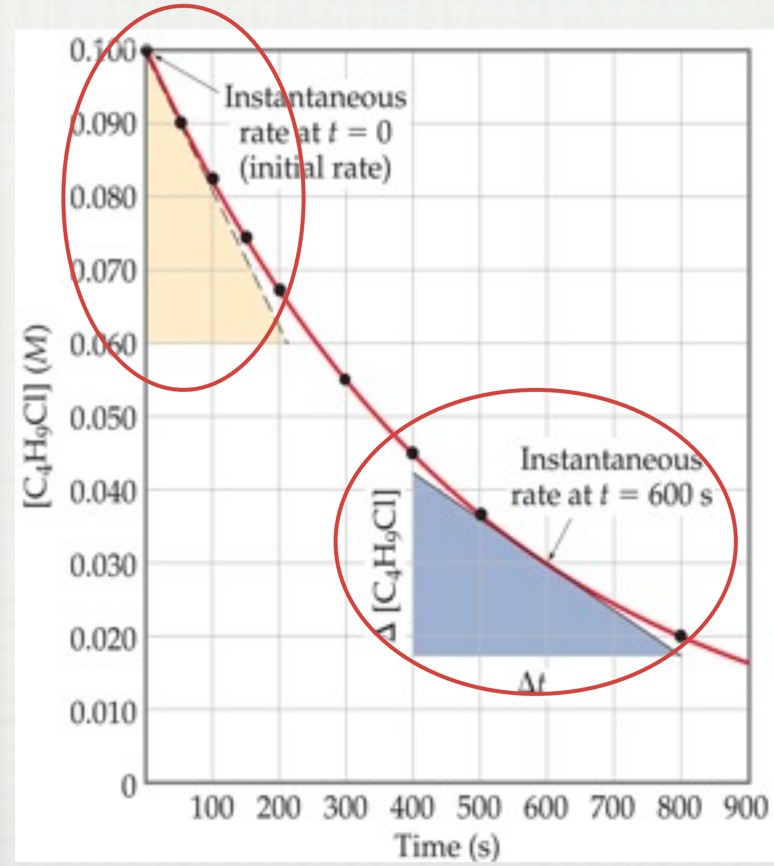
Time, t(s)	$[C_4H_9Cl]$ (M)	Average Rate (M/s)
0.0	0.1000	
		1.9×10^{-4}
50.0	0.0905	
		1.7×10^{-4}
100.0	0.0820	
		1.6×10^{-4}
150.0	0.0741	
		1.4×10^{-4}
200.0	0.0671	
		1.22×10^{-4}
300.0	0.0549	
		1.01×10^{-4}
400.0	0.0448	
		0.80×10^{-4}
500.0	0.0368	
		0.560×10^{-4}
800.0	0.0200	
10,000	0	

- Note that the average rate decreases as the reaction proceeds.
- This is because as the reaction goes forward, there are fewer collisions between the reacting molecules.

Reaction Rates and Stoichiometry

$C_4H_9Cl(aq) + H_2O(l) \longrightarrow C_4H_9OH(aq) + HCl(aq)$

- In this reaction, the ratio of C_4H_9Cl to C_4H_9OH is 1:1.
- Thus, the rate of disappearance of C_4H_9Cl is the same as the rate of appearance of C_4H_9OH.



$$\text{Rate} = \frac{-\Delta[C_4H_9Cl]}{\Delta t} = \frac{\Delta[C_4H_9OH]}{\Delta t}$$

Reaction Rates & Stoichiometry

Suppose that the mole ratio is *not* 1:1?

Example

$H_2(g) + I_2(g) \longrightarrow 2\,HI(g)$

2 moles of HI are produced for each mole of H_2 used.

$$rate = -\frac{\Delta [H_2]}{\Delta t} = \frac{1}{2}\frac{\Delta [HI]}{\Delta t}$$

The rate at which H_2 disappears is only half of the rate at which HI is generated

TRY IT!

$2\ N_2O_5 \rightarrow 4\ NO_2 + O_2$

NO_2 IS BEING PRODUCED AT A RATE OF $5.00\ X10^{-6}$ M/S. WHAT IS THE RATE OF DECOMPOSITION OF N_2O_5?

1) WRITE THE RATE EXPRESSION:

RATE = $\Delta\ [NO_2]\ /\ \Delta T\ = 5.00\ X10^{-6}$ M/S

2) LOOK AT THE RATIO IN THE EQUATION: FOR EVERY MOLE OF NO_2 MADE, 1/2 N_2O_5 IS DECOMPOSED

RATE = $-\ \Delta\ [N_2O_5]/\Delta T\ = 1/2\ \Delta[NO_2]\ /\Delta T$

SUB IN! $\Delta\ [N_2O_5]/\Delta T = 1/2\ (5.00\ X\ 10^{-6}$ M/S)

$= 2.5\ X\ 10^{-6}$ M/S

TRY IT!

- P.7 IN YOUR WORKBOOK
- #1-3, 9-11