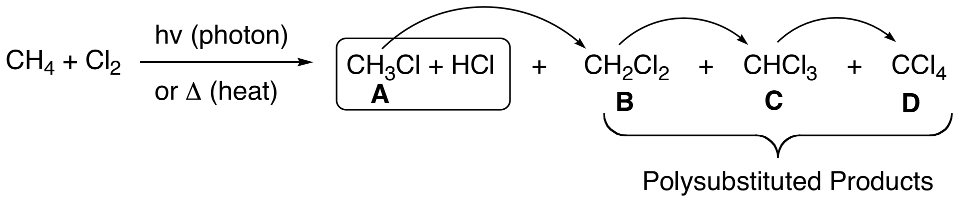
Ch. 4 The Study of Chemical Reactions

4.1 Three Factors in Every Reaction:

1. Mechanism: what is the step-by-step pathway by which old bonds break and new bonds form?
2. Thermodynamics: what are the energy changes, both for the overall reaction and for individual steps in the reaction mechanism?
3. Kinetics: How fast does a reaction occur? How do changes in reactant structure, reaction solvent, or reaction temperature speed up or slow down a reaction?

4.2 The Chlorination of Methane: A Case Study



Observations

-usually a mixture of products forms, including not only mono-chlorinated product **A**, but also polychlorinated products **B-D**.

1. Light (or heat) is required to initiate the reaction (energy required)
2. Blue light, absorbed by Cl_2, is most effective
3. High "quantum yield": one photon can result in conversion of thousands of methane reactant molecules into product molecules
    - Q: if light energy is needed, why isn't one photon needed for each reaction?

ANY MECHANISM MUST BE CONSISTENT WITH EXPERIMENTAL OBSERVATIONS

4.3 The Mechanism: Radical Chain Reaction

Balanced Reaction: $CH_4 + Cl_2 \xrightarrow{hv} CH_3Cl + HCl$

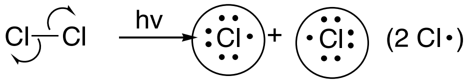
The mechanism must show all bonds broken and made:

Bonds Broken	Bonds Made
C-H	C-Cl
Cl-Cl	H-Cl

3 Phases in Mechanism

1. Initiation (gets it started)
2. **Propagation (keeps on going and going and going)**
3. Termination (what happens when it sometimes stops)

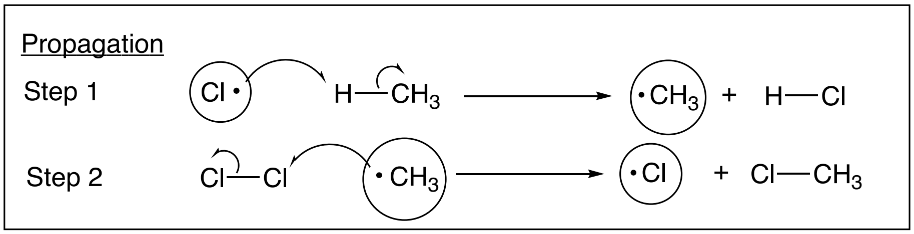
Initiation



"radical" something with an unpaired electron

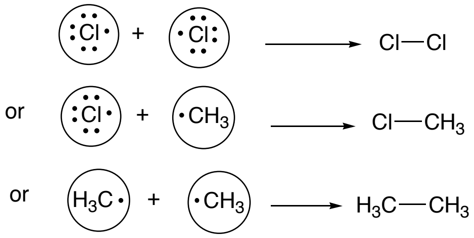
- In a radical initiation step, two reactive radicals form from a nonradical precursor

PROPAGATION



- In each propagation step, one reactive radicals reacts with a nonradical to produce a new reactive radical and a new nonradical.
- Since a reactive radical is reproduced in each step, you always have another reactive radical ready to keep the chain going.
- The chlorine radical produced in step two is the same as the chlorine radical that reacts in step 1.
- Thus you can sustain a repeating chain of step 1- step 2 -step 1- step 2 - step 1- step 2 - step 1- step 2 - step 1- step 2 - step 1- step 2 - etc.
- -As long as there is a radical around, the chain will keep going
- The sum of the two propagation steps is the overall balanced reaction

Termination



Mechanism Notes:

1. "Radical": Something with an unpaired electron.
   - Never satisfy octet rule→ highly unstable and always highly reactive.

2. Initiation is needed to produce a highly reactive radical. But once you've got one, initiation is done.

3. **The main action is the propagation phase. Memorize how that works.**

4. The propagation phase involves a repeating chain of events (step 1 – step 2 – step 1 – step 2 etc.) **"Chain reaction"**

5. The overall reaction is the sum of the two propagation steps
   - When you sum the propagation steps, notice that the methyl radical cancels itself out (what's formed in step one is erased in step 2) and the chlorine radical cancels itself out (what's formed in step two is erased in step 1).

6. Like initiation, termination occurs only occasionally

7. Notice:
   - Initiation: nonradical in → radicals out
   - Each Propagation Step: radical + nonradical → nonradical + radical
   - Any Termination Step: radical + radical → nonradical

4.4, 4.5 Free Energy, Enthalpy, Entropy

$$\Delta G = \Delta H - T\Delta S$$

ΔG: Free Energy: favorable reactions have negative ΔG
ΔH: Enthalpy: heat lost or gained
- ΔH<0 exothermic ΔH>0 endothermic

ΔS: Entropy: degree of randomness, disorder

In organic, enthalpy almost always dominates

Exothermic → Favorable **Endothermic → Unfavorable**

If you can figure out whether a reaction will be exothermic or not, you can tell whether it is energetically favorable or not.

- But, being energetically favorable still doesn't prove it will happen very fast… That's the kinetics issue, see later…

4.6 Bond Energies:

- **Exothermic reactions break weaker bonds and form stronger bonds**
- Exothermic steps (in a multistep reaction) also trade weaker for stronger
- Extensive tables of bond energies are available (Table 4.2) for when bonds break in half (to give two radicals)
- Often relative bond energies can be predicted by inspection

Bond Strength	Bond Energy (kcal/mol)	Molecule	Products	Radical Stability
Strongest	136	H—F →	H• + F•	Least stable
	103	H—Cl →	H• + Cl•	
	88	H—Br →	H• + Br•	
Weakest	71	H—I →	H• + I•	Most stable

Skills:

- Given bond energies, be able to rank bond strengths
- Given bond energies, be able to rank radical stabilities
- Given known radical stabilities, be able to predict relative bond strengths
- **Memorize the stability pattern for the halogen radicals**
- **Memorize the bond strength pattern for H-X bonds**
- **Memorize: C-X bonds have the same pattern: iodide is the weakest**

H_3C—F	H_3C—Cl	H_3C—Br	H_3C—I
109	84	70	56

- Just as acidity reflects anion stability, bond energy values reflect radical stability

Problems:

1. H_3C—SeH bonds are weaker than H_3C—OH bonds. Which is more stable, •SeH or •OH?

•Se is more stable.

2. Which is stronger, CH_3CH_2—Cl or CH_3CH_2—Br?

CH_3CH_2—Cl

Why are H-F and C-F bonds stronger than H-I and C-I bonds?

1. **Electronegativity and radical stability: (Remember)**
   - Radicals are short of octet rule → electron poor
   - The more electronegative fluorine is least willing to be electron poor. As you go down the table, electronegativity decreases and it's less problematic to become radical

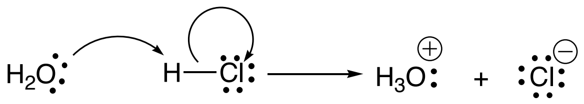
Problem: Rank the probable stability of the following radicals, 1 being most stable and 4 being least stable? (Use electronegativity to guide you…)

$H_3C\bullet$	$H_2N\bullet$	$HO\bullet$	$F\bullet$
1	2	3	4

2. Atomic size and orbital overlap:
   - Fluorine is small, and it's orbitals match up well size-wise with H and C resulting in strong overlap and strong bonds.
   - Iodine is big, so it's orbitals don't match up well or overlap so well with H or C resulting in weak bonds.

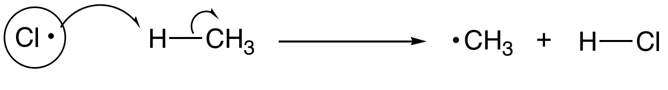
Two Types of Bond Breaking:

Heterolysis: one atom keeps both electrons (usual case)
- Ions are involved
- Arrow-pushing involves double-barbed arrows ()



Both electrons in the H-Cl bond went with the chlorine

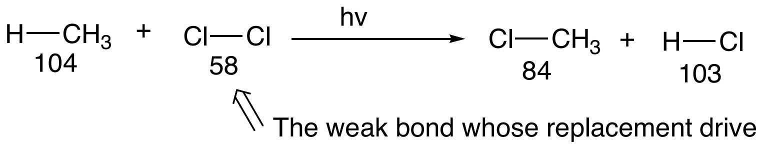
Homolysis: Bond breaks in half so that an electron goes with each atom (rare, but that's the type in this chapter)
- Radicals are involved
- Arrow-pushing involves single-barbed arrows ()



One electron in C-H bond goes off with carbon. The other stays with Hydrogen, and matches up with the electron from chlorine to make the new H-Cl bond.

4.7 Calculating Energy Changes

ΔH = (bond energies of bonds broken) - (bond energies of bonds formed)



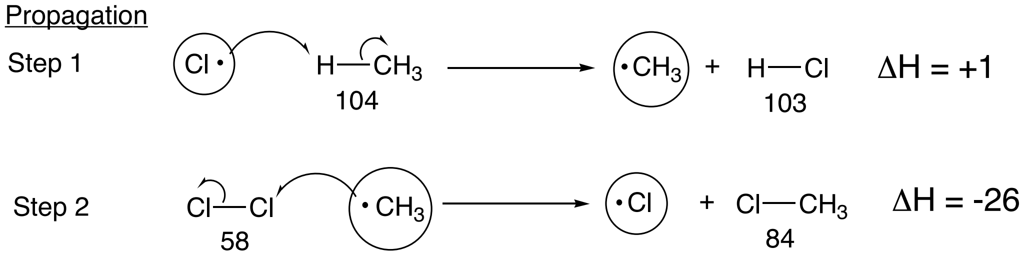
The weak bond whose replacement drives the reaction

Q1: What is ΔH: ΔH = (104 + 58) – (84 + 103) = -25 kcal/mol (exothermic)

Q2: Is the overall reaction energetically favorable? Yes

Notes:

- Compare the energies of the bonds broken versus the bonds made
- For an energetically **favorable** process, **<u>weaker bonds are replaced by stronger bonds</u>**
- With known bond energies, you can quantitatively calculate ΔH
- Even without bond energy numbers, a qualitative sense of bond strengths enables evaluation of whether or not a reaction makes sense energetically
- This type of analysis can be applied both to overall reactions, but also for individual steps in a multi-step reaction.



Q1: Which step is better? Step 2

Q2: Which step is likely to be the rate-limiting step? Step 1

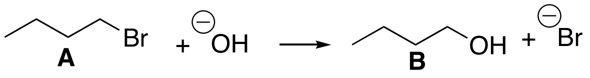
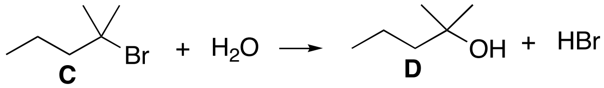
Q3: Note: Can you see what initiation would cost, and why a good chunk of energy is required to make it happen? 58. The 58kcal/mol is broken, while no new bond replaces it.

4.8 Kinetics, Reaction Rates, and Rate Laws

- Lots of reactions with seemingly favorable ΔH energetics don't happen very fast or at all
- We're often really interested in reaction speed ("kinetics"). Not so simple!
- Rate Law: relationship between reactant concentrations and overall rate

General: rate = $k[A]^x[B]^y$

- k is rate constant: each reaction has it's own unique rate constant.
- We will often be able to make qualitative predictions based on structural factors
- "x" and "y" are the "orders" of reactants A and B
- the "overall order" of a reaction = x + y

	overall rate law	overall order	individual orders
	$r = k[A]^1[^{\ominus}OH]^1$		
	$r = k[C]^1$		

- Different rate laws reflect different mechanisms
- Reactants that do not appear in a rate law do not appear in the mechanism until after the rate determining step
- The "k" values for the two reactions are **not** the same.
- Concentrations matter, for reactants that appear in the rate law
- Concentrations reflect not only how many moles of reactant are available, but also the amount of solvent.

Q1: If you use the same number of moles of reactants in reaction one, but you triple the volume of solvent, how will the rate change?

1/9 as fast. (Each reactant is diluted to 1/3 molarity. 1/3 x 1/3 = 1/9 as fast)

Q2: If you triple the volume of solvent for reaction two, again without changing the number of moles of reactants, how will it's rate change?

1/3 as fast. (The only reactant in the rate law is diluted to 1/3 molarity, so the rate will decrease proportionally.)

4.8 Activation Energies and Dependence of Rates on Temp

- So, if every reaction has it's own k value, what influences the "k" value?
- Arrhenius Equation: $k = Ae^{-Ea/RT}$
    - A is a constant
    - E_a or E_{act} is the "activation energy"
    - R is the ideal gas constant
    - T is the temperature
- Math: larger E_{act}/RT → smaller k (and slower reaction)
- Math: smaller E_{act}/RT → larger k (and faster reaction)

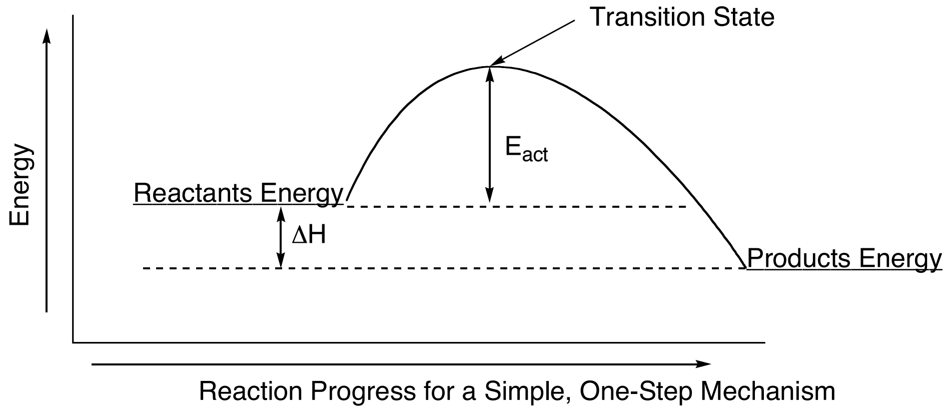
Practical Stuff

Temperature:

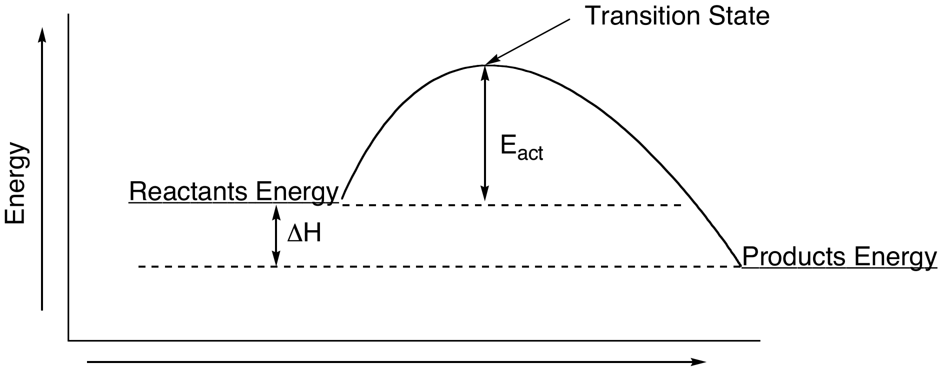
- Higher temp → higher k → faster reaction
- Lower temp → smaller k → slower reaction
- Crude guide: for every 10° rise in temp, the k value and reaction rate will double for an ordinary reaction. (This is super, super, super crude, though…)

Activation Energy: the minimum energy needed for a reaction

- It's the minimum energy required to cross the energy barrier between reactants and products
- The height of the barrier influences reaction speed.
- Activation barriers explain why many exothermic, energy-favorable reactions don't actually occur at room temperature
- Temperature reflects the average kinetic energy of the molecules; but some are always above average.
- **An increase in temperature can strongly increase the reaction rate because a small temperature increase can substantially increase the population of molecules with E_{act} (see Figure 4.2 on p. 140 for a nice picture of this).**

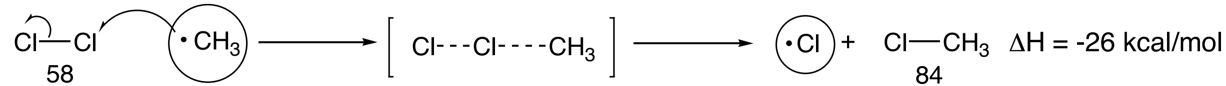


4.10 Transition States



Reaction Progress for a Simple, One-Step Mechanism

- The transition state is the **<u>highest, worst energy spot</u>** on the road from reactants to products
- The height of the transition state dictates the magnitude of the activation barrier (E_{act}). Thus the T-state has a huge impact on reaction rates.
- Why are T-states so much higher in energy than most products or reactants? And why do reaction with very favorable ΔH often have fairly high T-states?
    - o Because one full bond is better than two partial bonds. At the T-state, you are routinely at the transition between a breaking bond and a forming bond.



58

Weaker bonding than Product
Stronger bonding than transition state

T-State
Old Cl-Cl bond is partly broken
New Cl-CH_3 bond is partly formed
Combined energy: 54

84

Strongest Bonding

Since rates are affected by E_{act}, and E_{act}'s are determined by Transition States, → Transition states influence reactions rates.

- Lower transition state → faster reaction
- Higher transition state → slower reaction

<u>Transition-State Stability/Reactivity Principle</u>: The <u>more stable the transition state</u>, the <u>faster the reaction</u> will be.

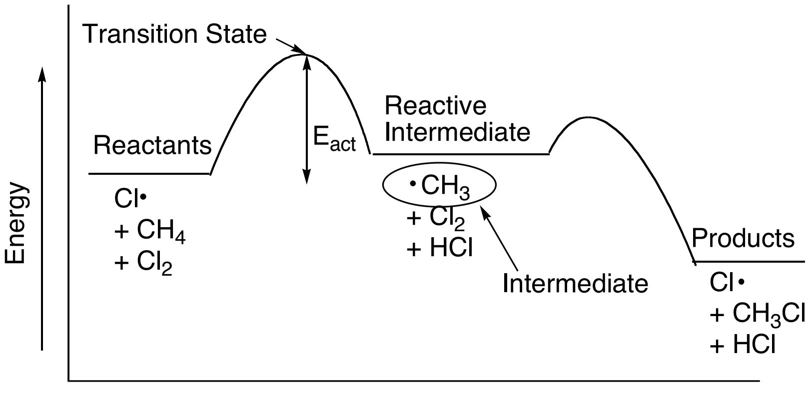
<u>Reactant Stability/Reactivity Principle</u>: The <u>more stable the reactant</u>, the <u>slower</u> it will react.

- A more stable reactant has lower starting energy. Therefore it has a larger E_{act} to get over the transition state.
- A less stable reactant has a higher starting energy, is closer to the T-state, and thus has a smaller energy barrier to cross.

<u>Product Stability/Reactivity Principle</u>: The <u>more stable the product</u>, the <u>faster the reaction</u>.

- A more stable product has lower energy. Often the T-state is stabilized/lowered by the same factors that may stabilize the products.

4.11 Rates of Multistep Reactions



Reaction Progress for a Simple, Two-Step Mechanism

- Most reactions involve 2 or more steps with one or more "intermediates" (•CH_3 is the key intermediate in the reaction above)
- An "**intermediate**" is something that forms temporarily, but then rapidly converts into something else. Normally the intermediate is highly reactive and has a very short lifetime. A significant population of intermediate never accumulates.
- The **transition state** for the overall reaction is still the **highest, worst energy spot** on the road from reactants to products
- There is only one transition state for the overall process, no matter how many steps
- The step that goes through the transition state will be the **slowest step** and is often referred to as **the rate-determining step** or the **slow step**.

Practical: To handle rates, identify and only think about the slowest step!!!

Practical: The rate determining step will **always** be the step leading to the worst, least stable intermediate. (•CH_3 < •Cl)

- Therefore the ability to recognize stability patterns for reactive intermediate radicals, cations, and anions is super useful

Product Stability/Reactivity Principle: The more stable the product, the faster the reaction.

- In multistep reactions, **the product that matters kinetically is the product of the rate-determining step. Which is often a reactive intermediate**.
- In order to apply the product stability/reactivity principle in multistep reactions, you'll need to:
  - Know the mechanism. (What is the rate determining step? And what kind of reactive intermediate is produced in that rate-determining step?)
  - Know how structural factors impact the relative stabilities of reactive intermediates. (For example, is a 3° radical better or worse than a 1° radical?)

Reactant Stability/Reactivity Principle: If a reactive intermediate is a reactant in a rate-determining step (rare, but this chapter), knowing stability patterns will help predict speeds.

4.12 Dependence of Halogenation Rates on Halogen

General reaction: $CH_4 + X_2 \rightarrow CH_3X + HX$

Rate determining step: $CH_4 + \bullet X \rightarrow \bullet CH_3 + HX$

Halogen	Rate Determining Step	E_{act} (kcal/mol)	•X Stability	•X Reactivity
F_2	$CH_4 + \bullet F \rightarrow \bullet CH_3 + HF$	1	Least	Most
Cl_2	$CH_4 + \bullet Cl \rightarrow \bullet CH_3 + HCl$	4		
Br_2	$CH_4 + \bullet Br \rightarrow \bullet CH_3 + HBr$	18		
I_2	$CH_4 + \bullet I \rightarrow \bullet CH_3 + HI$	34	Most	Least

- Iodine is not reactive enough; fluorine is actually too dangerous to use

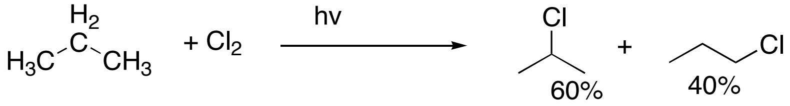
Reactant Stability/Reactivity Principle: The more stable the reactant, the slower it will react.

- The halogen radicals are reactants in the rate determining step.

4.13,14 Selective Halogenations of Higher Alkanes (Higher than Methane)

This is where most of the real problems will come from

A. Chlorination of Propane



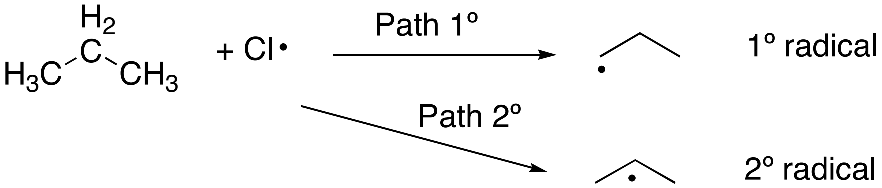
Notes

1. Two types of hydrogens leads to formation of two structural isomers
2. 2° hydrogens are more reactive than 1° hydrogens

Reactivity ratio: actual ratio is 3:2 or 1.5:1. Statistically expected ratio (two 2° hydrogens to six 1° hydrogens) would be 1:3. The ratio of actual (1.5) to the statistically expected (0.33) = 4.5. A secondary H is 4.5 times more reactive than a primary.

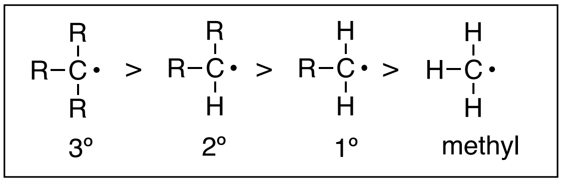
Why are 2° C-H's more reactive than 1° C-H's?

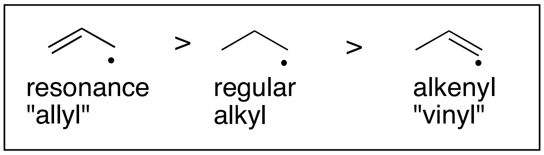
- Think rate determining step
- Think whether some stability/reactivity principle might apply



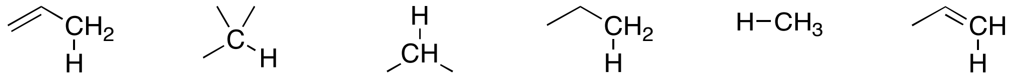
Path 2° is faster than path 1° because path 2° produces a more stable radical product. The path 2° transition-state is stabilized as a result. Product stability/reactivity principle.

B. Free Radical Stability Pattern: $3^o > 2^o > 1^o >$ methyl Memorize!



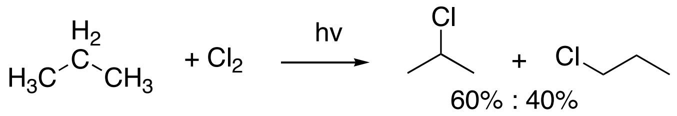


Resonance helps a lot (“**Allylic**”)
Being on an alkene is bad (“**Vinylic**”)

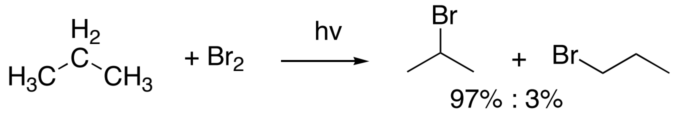


Bond Energy	87	91	95	98	104	111
Class	Allylic	3^o	2^o	1^o	Methyl	vinyl

C. Bromination of Propane



3:2 product ratio
4.5:1 selectivity for
2^o over 1^o hydrogens



33:1 product ratio
97:1 selectivity for
2^o over 1^o hydrogens

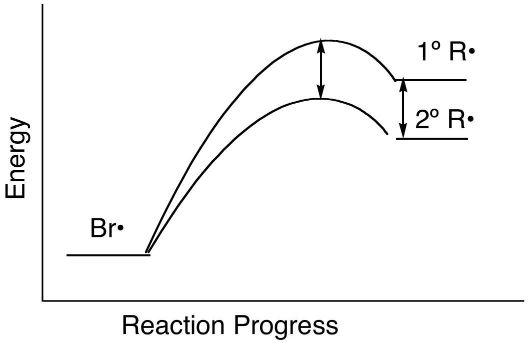
Notes
Bromine is **way** more selective than chlorine
Practical: to do a selective halogenation, use bromine rather than chlorine
Just as $2^o > 1^o$, so allylic $> 3^o > 2^o > 1^o >$ methyl > vinyl

D. Why bromination is so much more selective than chlorination:

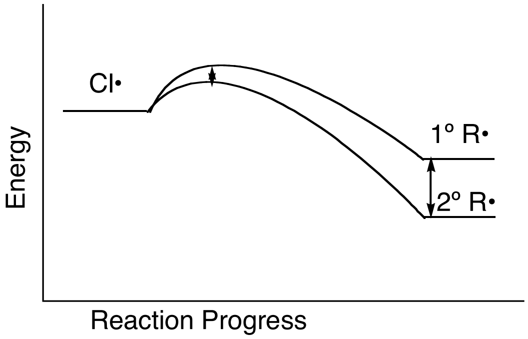
Reactant Stability/Reactivity/Selectivity Principle: Often a reaction can proceed to give either of two products, of unequal stability. The pathway leading to the more stable product will be preferred. However, the selectivity between formation of the more stable and less stable product will vary depending on the stability of the reactant. The more stable the reactant, the less reactive it will be and the more selective it will be. (The concept here is that a more stable reactant is less desperate to react, and is more choosy, better able to select the best possible pathway without using a less favorable pathway that would result in a less stable product. A more sophisticated picture is that a more stable reactant will have larger activation barriers to cross; it has a hard enough time crossing even the lowest transition state leading to the best possible product, and is much less likely to have the surplus energy required to cross the high transition state leading to the less stable product.)

Application to the Propane Halogenation Situation:

- Br• is more stable than Cl•,
- Therefore Br• is more selective and choosy to make the better 2° radical (leading to 2-bromopropane) rather than the less stable 1° radical (leading to 1-bromopropane).
- Cl• is less stable, and really wants to react. So it doesn't wait around for a weak 2° hydrogen; it often settles for a stronger 1° hydrogen even though it gives an inferior1° radical product (in the rate determining step)
- "Beggars can't be choosers": the less stable, more reactive Cl• is the "beggar" than can't be as choosy as the more stable, less reactive Br•



- Endothermic
- Late transition states
- T-states resemble product(s)
- The energy gap between alternate T-states is almost as large as the energy gap between alternate products.
- The strong energy difference between the two T-states results in high selectivity



- Exothermic
- Early transition states
- T-states resemble reactant
- The energy gap between alternate T-states isn't nearly as large as the energy gap between alternate products.
- The limited energy difference between the two T-states results in limited selectivity

E. Hammond Postulate:

- For an exothermic step, the T-state is "early" and much resembles reactants
- For an endothermic step, the T-state is "late" and much resembles products

Who Cares?

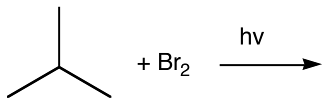
- Transition-states are important! The Hammond postulate helps us to understand what the structure of the transition state is like.
- For most multi-step reactions, the transition-states are late and product-like.
- **For late, product-like transition states, assume that any structural factor that stabilizes the product (of the rate-determining step) will also stabilize the transition state and increase the reaction rate. (This is the basis of the product stability/reactivity principle.)**
- KEY: Remember that you must always be thinking about the products of the rate-determining step, which will routinely be a reactive intermediate that does not appear as a product in the balanced reaction.

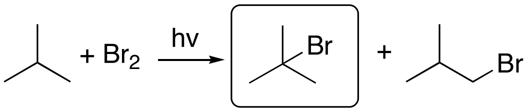
Alkane Brominations:

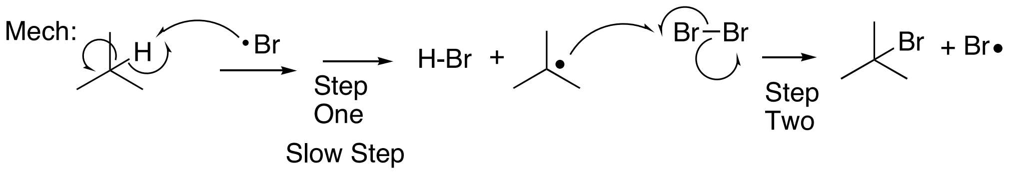
Skills:

1. Identify all possible monosubstituted products
2. Identify the Major Product
   - Consider all possible radicals. The carbon that gives the most stable radical will be the carbon that gets brominated preferentially.
   - This is true because the rate determining step is the step in which a hydrogen is abstracted and a carbon radical is formed.
   - Thus, according to the product stability/reactivity principle, the pathway leading via the best carbon radical is the preferred path.
3. Write the mechanism for chain propagation (with detailed arrows)

1. Do all three things for:







2. Identify the Major Product for each of the following: