Heat and Related Properties:

Enthalpy

&

Heat Capacities
Enthalpy $\equiv H \equiv U + PV$  

Greek: \textit{en} + \textit{thalpein} = “to heat in”

Sometimes “H” is referred to as “the heat content” of the system – I don’t like it!!

**Significance**

\[ H = f(\text{state functions}) \quad \therefore \quad H \text{ is a state function:} \]

\[
\int_1^2 dH = H_2 - H_1 = \Delta H \quad \text{and} \quad \int dH = 0
\]

**Significance of $dH$ and of $\Delta H$:**

\[
dH = dU + \underbrace{d(P_{\text{in}}V)}_{[p_{\text{ex}}=P_{\text{in}}]} = dU + p_{\text{ex}}dV = dq + \underbrace{dw}_{p_{\text{ex}}dV} + p_{\text{ex}}dV = dq
\]

\[dH = dq_P \quad \Delta H = q_P \quad (\text{for reversible PV-work only})\]

\[dU = dq_V \quad \Delta U = q_V \quad (\text{for PV-work only})\]

The change in enthalpy (or internal energy) is the energy transferred as heat at constant $P$ (or $V$).
Heating @ [P], @ [V]

\[ dH = dq_P \]
\[ \Delta H = q_P \]  \hspace{1cm} \text{(reversible PV-work)}

\[ dU = dq_V \]
\[ \Delta U = q_V \]  \hspace{1cm} \text{(PV-work)}

\( q_P = \text{Heating at [P]}: \)
\[ T \uparrow, \ V \neq \text{const.} \]

\[ q_V = \text{Heating at [V]}: \]
\[ T \uparrow, \ P \neq \text{const.} \]

\( q_P \) vs. \( q_V \) \hspace{1cm} \text{(which is greater?)}

Difference between \( q_P \) and \( q_V \) is the difference between \( \Delta H \) and \( \Delta U \):

\[ H \equiv U + PV \Rightarrow dH = dU + d(PV) \Rightarrow \Delta H = \Delta U + \int d(PV) \]
\[ dq_P = dq_V + d(PV) \Rightarrow q_P = q_V + \int d(PV) \]

\[ \int d(PV) = (PV)_2 - (PV)_1 = \Delta(PV). \text{ For an ideal gas: } \Delta(PV) = nR\Delta T \]

solid: \( \sim 0, \ [P,V \equiv \text{const.}] \)
Amount of heat needed to raise a certain amount of substance by 1°

\[ C_s \equiv \text{specific heat capacity,} \quad \frac{\text{cal or J}}{\text{g} \cdot \text{deg}} \]

\[ \bar{C} \equiv \text{molar heat capacity,} \quad \frac{\text{cal or J}}{\text{mol} \cdot \text{deg}} \]

For example:

For \( \text{H}_2\text{O}(l) \): \( C_s \approx 1 \text{ cal/g} \cdot \text{deg} \) \( \Rightarrow \) \( \bar{C} = \) _____________

Is “C” a state function?
\[ C_P = \frac{dq_P}{dT}, \quad dq_P = dH \]  
\[ (rev\ PV\text{-work}) \]

\[ C_V = \frac{dq_V}{dT}, \quad dq_V = dU \]  
\[ (PV\text{-work}) \]

Which is larger for a particular substance, \( C_P \) or \( C_V \)?

What about for solids?

\( C_P \) is directly dependent on the expansion ability of the substance:

Coefficient of thermal expansion

Thermal expansivity

Cubic expansion coefficient

\[ \alpha \equiv \frac{1}{V} \left( \frac{\partial V}{\partial T} \right)_P \]  
\( (Note\ each\ factor) \)

For an ideal gas, \( \alpha = 1/T \)  
(Do it!)
\[ C_V = \left( \frac{\partial U}{\partial T} \right)_V \]
For any Pure Substance

\[ C_v = \left( \frac{\partial U}{\partial T} \right)_v \]

For an Ideal gas:

\[ C_v = \frac{dU}{dT} \]

\[ C_p = \left( \frac{\partial H}{\partial T} \right)_p \]

\[ C_p = \frac{dH}{dT} \]

For an Ideal gas:

\[ C_p = \frac{dH}{dT} = \frac{d(U + PV)}{dT} = \frac{dU}{dT} + \frac{d(PV)}{dT} = C_v + nR \]

\[ \overline{C_p} = \overline{C_v} + R \]
Notes:
The more complex the molecule,
1. the greater its $C_P$, and
2. the greater the increase with rising $T$. 
\( \bar{C}_P \) as a function of Temperature from 300 to 1500 \( K \): \( \bar{C}_P = a + bT + cT^2 \)

<table>
<thead>
<tr>
<th>Molecule ((g))</th>
<th>(a / JK^{-1}mol^{-1})</th>
<th>(10^3b / JK^{-2}mol^{-1})</th>
<th>(10^7c / JK^{-3}mol^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>(H_2)</td>
<td>29.066</td>
<td>-0.836</td>
<td>20.113</td>
</tr>
<tr>
<td>(O_2)</td>
<td>25.503</td>
<td>13.612</td>
<td>-42.553</td>
</tr>
<tr>
<td>(Cl_2)</td>
<td>31.696</td>
<td>10.144</td>
<td>-40.375</td>
</tr>
<tr>
<td>(N_2)</td>
<td>26.984</td>
<td>5.910</td>
<td>-3.376</td>
</tr>
<tr>
<td>(HCl)</td>
<td>28.166</td>
<td>1.809</td>
<td>15.465</td>
</tr>
<tr>
<td>(H_2O)</td>
<td>30.206</td>
<td>9.936</td>
<td>11.14</td>
</tr>
<tr>
<td>(CO_2)</td>
<td>26.648</td>
<td>42.262</td>
<td>-142.4</td>
</tr>
<tr>
<td>(CH_4)</td>
<td>14.143</td>
<td>75.495</td>
<td>-179.64</td>
</tr>
<tr>
<td>(C_2H_6)</td>
<td>9.404</td>
<td>159.836</td>
<td>-462.28</td>
</tr>
</tbody>
</table>

(many other tables for solids, liquids, gases)

Note: In some tables, \( C_P = a + bT + cT^2 \).
\[ C_P^0 \text{ of } \text{SO}_2 \text{ vs. } T \]

\[
\left( \frac{\partial S}{\partial T} \right)_P = \frac{C_P}{T}
\]

(Alberty, p. 98)
\[
\left( \frac{\partial S}{\partial T} \right)_P = \frac{C_p}{T}
\]
Molecular Interpretation of $C_V$ for an Ideal Gas

Molecular Energies Associated With Atomic (Nuclear) Motions

Each atom in a molecule can move independently in 3 basic directions, i.e., each movement can be described by a vector:

$\therefore$ # of degrees of freedom for the movement of each atom in a molecule = 3

For N atoms in a molecule: Total # of degrees of freedom = 3N

But this seemingly chaotic atomic motion that is executed simultaneously by all of the atoms can be dissected into a set of ordered motions: Translation, Rotation, and Vibration.

$\text{Heat (thermal energy) mainly influences translational, rotational, and vibrational motions and not the high-lying electronic energy levels (}1\text{ eV} = 96,500\text{ J/mol}).$

$E_{\text{thermal}} \propto kT \text{ (per atom: KMT).}$

$RT \sim 2500\text{ J/mol, }298\text{ K}$

$E_{\text{vib},6} \sim 2640\text{ J/mol (}4000\text{ cm}^{-1})}$

$\text{Vib: } E_1 = \frac{1}{2} h \nu N_A \sim 720-72\text{ J/mol}$

$\text{Vib: } E_2 = \frac{1}{2} h \nu N_A \sim 240-24\text{ J/mol, wavenumber } \equiv \tilde{\nu} = 4000 - 400\text{ cm}^{-1}$

$\text{Tr}$

$\text{Rot}$
Total # of Degrees of Freedom for a molecule with N atoms = 3N

**Translation:**
Molecule moves as if “frozen”; movement of the center of mass as a vector.
# of independent translational motions (along 3 axes) = 3

**Rotation:**
Whole molecule twists as a “frozen” block about axes of symmetry, without moving the center of mass.
For **non-linear** molecules, # of independent rotations (about 3 axes) = 3
For **linear** molecules, # of independent rotations (about 2 axes only) = 2
(No energy is needed to “spin” a linear molecule about the internuclear axis, because the atoms (as point masses) do not really move; no movement, no moment of inertia, I.)

**Vibrations (the remaining motions):**
Each atom can “twist & shout”:
Bonds are like “springs” and bond angles are flexible; bonds can stretch and contract, bonds can bend, etc., without moving the center of mass, and without rotating the molecule. So, here’s what’s left of the total 3N motions:
For **non-linear** molecules, # of independent motions = 3N–6
For **linear** molecules, # of independent motions = 3N–5

Examples for H₂O & CO₂ (continued)
Energies of Independent Motions for an Ideal Gas: Principle of Equipartition of Energy

Recall the results of the Kinetic Molecular Theory for an Ideal Gas:

**K.E. of translation** = \( \frac{3}{2} kT \) or \( \frac{3}{2} RT = 3(\frac{1}{2})RT \)

**Axiom 1:** Energy of each independent translational direction = \( \frac{1}{2}RT \).
(The total translational energy has been **equally partitioned** into 3 parts.)

**Axiom 2:** Energy of any independent motion (rotation or vibration) = \( \frac{1}{2}RT \).

**Note:**
Potential Energy of translation or rotation = 0
**Intramolecular** P.E. associated with every vibrational mode ≠ 0

Recall: The bond is a “spring” and a **restoring force** acts on it as it is stretching:

**Hooke’s Law:** \( F = -kx \), \( k \equiv \) force (or spring) constant \( \Rightarrow \) \( F = -d\phi/dr \), \( \phi = \text{P.E.} \)

<table>
<thead>
<tr>
<th>Motion</th>
<th>Molecule</th>
<th>KE</th>
<th>PE</th>
<th>U (total)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Translation</td>
<td>(any)</td>
<td>( 3 \cdot (\frac{1}{2}RT) )</td>
<td>0</td>
<td>( 3 \cdot (\frac{1}{2}RT) )</td>
</tr>
<tr>
<td>Rotation</td>
<td>Non-linear</td>
<td>( 3 \cdot (\frac{1}{2}RT) )</td>
<td>0</td>
<td>( 3 \cdot (\frac{1}{2}RT) )</td>
</tr>
<tr>
<td></td>
<td>Linear</td>
<td>( 2 \cdot (\frac{1}{2}RT) )</td>
<td></td>
<td>( 2 \cdot (\frac{1}{2}RT) )</td>
</tr>
<tr>
<td>Vibration</td>
<td>Non-linear</td>
<td>((3N-6) \cdot (\frac{1}{2}RT))</td>
<td>((3N-6) \cdot (\frac{1}{2}RT))</td>
<td>(2 \cdot (3N-6) \cdot (\frac{1}{2}RT))</td>
</tr>
<tr>
<td></td>
<td>Linear</td>
<td>((3N-5) \cdot (\frac{1}{2}RT))</td>
<td>((3N-5) \cdot (\frac{1}{2}RT))</td>
<td>(2 \cdot (3N-5) \cdot (\frac{1}{2}RT))</td>
</tr>
</tbody>
</table>

\[ \therefore \text{For an ideal gas: } C_V = dU/dT \quad \& \quad C_P - C_V = R \]
## Theoretical (Classical) vs. Experimental Heat Capacities

<table>
<thead>
<tr>
<th>Molecule</th>
<th>$U_{\text{translation}}/RT$</th>
<th>$U_{\text{rotational}}/RT$</th>
<th>$U_{\text{vibrational}}/RT$</th>
<th>$U_{\text{total}}/RT$</th>
<th>$\overline{C_V}/R$</th>
<th>$\overline{C_P}/R$</th>
<th>Experimental $\overline{C_P}/R$</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>1.5</td>
<td>0</td>
<td>0</td>
<td>1.5</td>
<td>1.5</td>
<td>2.5</td>
<td>2.500</td>
</tr>
<tr>
<td>H$_2$</td>
<td>1.5</td>
<td>1</td>
<td>1</td>
<td>3.5</td>
<td>3.5</td>
<td>4.5</td>
<td>3.468</td>
</tr>
<tr>
<td>O$_2$</td>
<td>1.5</td>
<td>1</td>
<td>1</td>
<td>3.5</td>
<td>3.5</td>
<td>4.5</td>
<td>3.533</td>
</tr>
<tr>
<td>I$_2$</td>
<td>1.5</td>
<td>1</td>
<td>1</td>
<td>3.5</td>
<td>3.5</td>
<td>4.5</td>
<td>4.435</td>
</tr>
<tr>
<td>CO$_2$</td>
<td>1.5</td>
<td>1</td>
<td>4</td>
<td>6.5</td>
<td>6.5</td>
<td>7.5</td>
<td>4.466</td>
</tr>
<tr>
<td>H$_2$O</td>
<td>1.5</td>
<td>1.5</td>
<td>3</td>
<td>6</td>
<td>6</td>
<td>7</td>
<td>4.038</td>
</tr>
<tr>
<td>NH$_3$</td>
<td>1.5</td>
<td>1.5</td>
<td>6</td>
<td>9</td>
<td>9</td>
<td>10</td>
<td>4.285</td>
</tr>
<tr>
<td>CH$_4$</td>
<td>1.5</td>
<td>1.5</td>
<td>9</td>
<td>12</td>
<td>12</td>
<td>13</td>
<td>4.286</td>
</tr>
</tbody>
</table>

1. Heat mainly influences translational, rotational, and vibrational motions and **not** the high-lying electronic energy levels.
2. The classically calculated “C” assumes that **all** molecular motions contribute **equally** at **all** temperatures.
3. Experimental C’s: At low T, vibrational motions are negligible or small: $C_{\text{exp}} < C_{\text{class}}$.
   At high T, vibrations are important: $C_{\text{exp}} \sim C_{\text{class}}$.
4. As $T \rightarrow \infty$, real gas $\rightarrow$ ideal gas.

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Heat Capacities of Atomic Solids

At Room and Higher Temps: Law of Dulong et Petit (1819)
For metallic elements, $C \sim 3R = 6 \text{ cal/mol}\cdot\text{deg}$ (actually $6 \pm 0.3$)

Consider an atom bound in an atomic crystal:

Each atom oscillates (with $\nu$) about equilibrium position in 3 independent directions (no translations, no rotations):

$U_{\text{vib}} = \text{KE} + \text{PE} = 3(\frac{1}{2}RT) + 3(\frac{1}{2}RT) = 3RT$ (for a mole)

$C_V = \frac{dU}{dT} = 3R$

At Low Temps: Debye Theory (1912):

$\bar{C}_V \propto T^3$

Heat capacities of all atomic solids should lie on the same curve when $C_V$ is plotted versus $T/\theta_D$:

$\theta_D \equiv \text{Debye temperature} \equiv \hbar \nu/k_B$,

$\nu \equiv \text{frequency of oscillation} \equiv f(m,k)$,

$k \equiv \text{force constant}$

At low $T$ ($\leq 15 \text{ K}$): $C_V = a(T/\theta_D)^3$. 

Calculations: \( \Delta U = q_V \) from \( C_V \) & \( \Delta H = q_P \) from \( C_P \)

\[
\bar{C}_V = \frac{d\bar{q}_V}{dT} = \left( \frac{\partial \bar{U}}{\partial T} \right)_V
\]

\[dU = dq_V = n\bar{C}_VdT \Rightarrow \Delta U = q_V = n\int \bar{C}_VdT \quad (PV\text{-work only})\]

In general: \( \bar{C}_V = f(T) = a + bT + cT^2 + \cdots \) (or something similar)

If \( \bar{C}_V \approx \text{constant} \) \( \Rightarrow \) \( \Delta U = q_V \approx n\bar{C}_V\Delta T \) or \( mC_V\Delta T, [V, C_V] \)

\[
\bar{C}_P = \frac{d\bar{q}_P}{dT} = \left( \frac{\partial \bar{H}}{\partial T} \right)_P
\]

\[dH = dq_P = n\bar{C}_PdT \Rightarrow \Delta H = q_P = n\int \bar{C}_PdT \quad (\text{rev } PV\text{-work only})\]

In general: \( \bar{C}_P = f(T) = a' + b'T + c'T^2 + \cdots \) (or something similar)

If \( \bar{C}_P \approx \text{constant} \) \( \Rightarrow \) \( \Delta H = q_P \approx n\bar{C}_P\Delta T \) or \( mC_P\Delta T, [P, C_P] \)
Degrees of Freedom or Number of Independent Variables

Every substance has an **Equation of State**, which relates the **State Variables** \((P,V,T,n)\) to each other:

\[ \therefore \text{Total # of degrees of freedom} = 3 \]

But, **for a given quantity “n”**, \# of degrees of freedom \(= 2\)

Thermodynamic Properties are State Functions that are expressed in terms of the **independent** state variables.

For example:

\[
U = U(P,T) \quad or \quad U = U(V,T) \quad or \quad U = U(P,V)
\]

\[
H = H(P,T) \quad or \quad H = H(V,T) \quad or \quad H = H(P,V)
\]

(continued)
\[ U = U(V, T): \quad \frac{dU}{dT} \bigg|_V \, dT + \frac{dU}{dV} \bigg|_T \, dV = C_v \, dT + \pi_T \, dV \]

\[ H = H(P, T): \quad \frac{dH}{dT} \bigg|_P \, dT + \frac{dH}{dP} \bigg|_T \, dP = C_p \, dT + \frac{dH}{dP} \bigg|_T \, dP \]

Internal pressure coefficient \( \equiv \pi_T \equiv \left( \frac{\partial U}{\partial V} \right)_T \) (note the units)

**For an ideal gas**: \( U = U(T) \) only! (Recall the Kinetic Molecular Theory)

\[ \therefore \pi_T = \left( \frac{\partial U}{\partial V} \right)_T = 0 \quad \Rightarrow \quad dU = n\bar{C}_V \, dT \]

\[ H = H(T) \) only! Proof: \( H \equiv U + PV = U(T) + nRT = H(T) \) only!

\[ \therefore \left( \frac{\partial H}{\partial P} \right)_T = 0 \quad \Rightarrow \quad dH = n\bar{C}_P \, dT \]

\[ \therefore \text{For an isothermal process involving an ideal gas:} \quad \Delta H = \Delta U = 0 \]

\[ \begin{align*}
  dU &= n\bar{C}_V \, dT \\
  dH &= n\bar{C}_P \, dT
\end{align*} \]

For any pure substance @ [V]

\[ \begin{align*}
  dU &= n\bar{C}_V \, dT \\
  dH &= n\bar{C}_P \, dT
\end{align*} \]

For any pure substance @ [P]

For an ideal gas

**First Law Problems**: heat capacity: 9, 11, 12, 66.
internal pressure coefficient $t \equiv \pi_T \equiv \left( \frac{\partial U}{\partial V} \right)_T$
internal pressure coefficient $t \equiv \pi_T \equiv \left( \frac{\partial U}{\partial V} \right)_T$

$\pi_T$ is a measure of the intermolecular forces

$\pi_T < 0$  
$(-slopes)$

$\pi_T > 0$  
$(+slopes)$

(for stable systems)
We often encounter the following types of PD’s:

\[ \left( \frac{\partial U}{\partial T} \right)_P \quad \text{and} \quad \left( \frac{\partial U}{\partial T} \right)_V \]

How do we relate one to the other?
For example, how do we relate the first in terms of the second?

The “Trick” is to “start with the 2\textsuperscript{nd} and create the 1\textsuperscript{st}”:

1. Begin with the total differential equation that includes the second PD:

\[ U = U(T,V) \quad \Rightarrow \quad dU = \left( \frac{\partial U}{\partial T} \right)_V dT + \left( \frac{\partial U}{\partial V} \right)_T dV \]

2. Next, CREATE the first PD from this eqn by dividing by “dT”, and impose the condition of constant P on each new PD:

\[ \left( \frac{\partial U}{\partial T} \right)_P = \left( \frac{\partial U}{\partial T} \right)_V + \left( \frac{\partial U}{\partial V} \right)_T \left( \frac{\partial V}{\partial T} \right)_P \]
Relationship Between $C_P$ and $C_V$

\[
C_P - C_V = \left( \frac{\partial H}{\partial T} \right)_P - \left( \frac{\partial U}{\partial T} \right)_V
\]

\[
H = U + PV
\]

\[
= \left( \frac{\partial U}{\partial T} \right)_P + P \left( \frac{\partial V}{\partial T} \right)_P - \left( \frac{\partial U}{\partial T} \right)_V
\]

\[
\text{similar partial derivatives}
\]

From before, for similar PD’s:

\[
\left( \frac{\partial U}{\partial T} \right)_P = \left( \frac{\partial U}{\partial T} \right)_V + \left( \frac{\partial U}{\partial V} \right)_T \left( \frac{\partial V}{\partial T} \right)_P
\]

\[
= P \left( \frac{\partial V}{\partial T} \right)_P + \left( \frac{\partial U}{\partial V} \right)_T \left( \frac{\partial V}{\partial T} \right)_P
\]

\[
\text{work produced per unit increase in Temp.}
\]

\[
\text{energy per unit Temp required to separate the molecules against intermolecular attraction}
\]

(continued)
\[ C_P - C_V = \left[ P + \left( \frac{\partial U}{\partial V} \right)_T \left( \frac{\partial V}{\partial T} \right)_P \right] + + + > 0 \]

(for stable sys.; see before)

coefficient of thermal expansion \( \equiv \alpha \equiv \frac{1}{V} \left( \frac{\partial V}{\partial T} \right)_P \)

internal pressure coefficient \( \equiv \pi_T \equiv \left( \frac{\partial U}{\partial V} \right)_T \)

\[ C_P - C_V = \alpha (P + \pi_T) V \]

(for any pure substance)

For an ideal gas: \[ C_P - C_V = nR \] \( \Rightarrow \) \[ \bar{C}_P - \bar{C}_V = R \]

\{ extra energy required to heat a mole of ideal gas by \( 1^0 \) at const \( P \)
over that required to heat it at const \( V \) \}

For a solid: \[ \bar{C}_P \cong \bar{C}_V \]

(continued)
\[ C_P - C_V = \alpha (P + \pi_T) V \]  
(for any pure substance)

\[ \pi_T \equiv \left( \frac{\partial U}{\partial V} \right)_T = T \left( \frac{\partial P}{\partial T} \right)_V - P \]  
(proved later, from S)

\[ \left( \frac{\partial P}{\partial T} \right)_V = \frac{\alpha}{\kappa} \]  
(proved later, from Euler’s Cycle Rule)

Calculate \( \pi_T \) for van der Waals gas and for another real gas.

Isothermal compressibility  
\[ \equiv \kappa \text{ or } \beta \equiv -\frac{1}{V} \left( \frac{\partial V}{\partial P} \right)_T \]  
(“kappa”)  
engineering

Coefficient of thermal expansion  
\[ \equiv \alpha \equiv \frac{1}{V} \left( \frac{\partial V}{\partial T} \right)_P \]

Calculate \( \alpha \) and \( \kappa \) for some real gases and compare with an ideal gas.

\[ C_P - C_V = \left( \frac{\alpha^2}{\kappa} \right) TV \]  
Thus, especially for liquids, it’s not enough to look just at \( \alpha \), but we need to look at \( \alpha^2/\kappa \).

Examples appear on next slide.

(continued)
\[ \overline{C}_P - \overline{C}_V = \left( \alpha^2 / \kappa \right) T V \]

Coefficient of thermal expansion \( \equiv \alpha \equiv \frac{1}{V} \left( \frac{\partial V}{\partial T} \right)_p \)

Isothermal compressibility \( \equiv \kappa \equiv -\frac{1}{V} \left( \frac{\partial V}{\partial P} \right)_T \)

<table>
<thead>
<tr>
<th>Substance</th>
<th>(10^4 \alpha / K^{-1})</th>
<th>(10^6 \kappa / atm^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>2.1</td>
<td>49.6</td>
</tr>
<tr>
<td>Benzene</td>
<td>12.4</td>
<td>90.5</td>
</tr>
<tr>
<td>Lead</td>
<td>0.861</td>
<td>2.21</td>
</tr>
<tr>
<td>Diamond</td>
<td>0.030</td>
<td>0.187</td>
</tr>
</tbody>
</table>

**Examples:**

For water at \(25^0C\) : \( \overline{C}_V = 74.8 \ J/K \cdot \text{mol} \). Calculate its \( \overline{C}_P \). (Answer: 75.3 \ J/K \cdot \text{mol})

For benzene at \(25^0C\) : \( \overline{C}_P = 134 \ J/K \cdot \text{mol} \). Calculate its \( \overline{C}_V \). (Answer: 89 \ J/K \cdot \text{mol})

(Note: \(d = 0.879 \ g/mL\))

**First Law Problems:** heat capacity: 26-31.