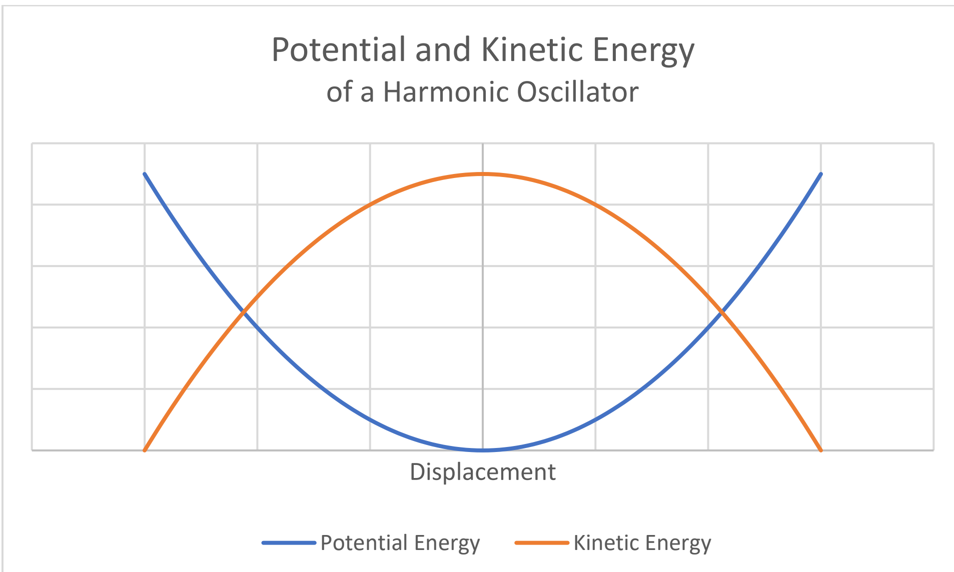
1. Convert the temperatures indicated to complete the following table

ºF	ºC	K
77	25	298.15
98.6	37	310.15
212	100	373.15
-40	-40	233.15
32	0	273.15

2. Make a graph representing the potential energy of a harmonic oscillator as a function of displacement from equilibrium. On the same graph, include a function describing the kinetic energy as a function of displacement from equilibrium as well as the total energy of the system.



3. Calculate the work required to move a 3.2 kg mass 10.0 m against a resistive force of 9.80 N.

Work is given by the negative product of force and displacement.

$$w = -F\,\Delta x$$

So,

$$w = -(9.80\ N)(10.0\ m)\left(\frac{J}{N\ m}\right) = -98.0\ J$$

4. Calculate the work needed for a 22.4 L sample of gas to expand to 44.8 L against a constant external pressure of 0.500 atm.

The work of expansion of a gas against a constant external pressure is

$$w = -p_{ext}\Delta V$$

So

$$w = -(0.500\ atm)(44.8\ L - 22.4\ L) = -11.2\ atm\ L$$

But what is an *atm L*? This can be converted to J using the ratio of the gaslaw constant expressed in appropriate units:

$$-(11.2\ atm\ L)\left(\frac{8.314\frac{J}{mol\ K}}{0.08206\frac{atm\ L}{mol\ K}}\right) = -11347.4\ J = -11.3\ kJ$$

Alternatively, one can convert to MKS units by recognizing that

$$1\ atm = 101325\ Pa$$
$$1000\ L = 1\ m^3$$
$$1\ Pa\ m^3 = 1\ J$$

So

$$-11.2\ atm\ L \cdot \frac{101325\ Pa}{1\ atm} \cdot \frac{m^3}{1000\ L} \cdot \frac{J}{Pa\ m^3} \cdot \frac{kJ}{1000\ J} = -11.3\ kJ$$

5. If the internal and external pressure of an expanding gas are equal at all points along the entire expansion pathway, the expansion is called “reversible.” Calculate the work of a reversible expansion for 1.00 mol of an ideal gas expanding from 22.4 L at 273 K to a final volume of 44.8 L.

The work of a reversible expansion (expressed in differential form) is given by

$$dw = -p\ dV$$

So, to get *w*, one must integrate!

$$w = -\int_{V_1}^{V_2} p\ dV$$

But pressure is very much dependent on volume for an ideal gas.

$$p = \frac{nRT}{V}$$

So

$$w = -nRT\int_{V_1}^{V_2}\frac{dV}{V}$$

For an isothermal, reversible expansion. (If the gas was not ideal, a difference equation of state might be used to substitute for *p* in terms of *V*, *n*, *R*, and *T*.) After integration, the expression becomes

$$w = -nRT\ln\left(\frac{V_2}{V_1}\right)$$

Substituting the values from the problem,

$$w = -(1.00\ mol)\left(8.314\frac{J}{mol\ K}\right)(273\ K)\ln\left(\frac{44.8\ L}{22.4\ L}\right)$$
$$= -15\underline{7}3.3\ J$$

6. The reaction of combustion of the gasoline (assuming it is composed of n-octane only) is given by the following chemical equation:

$$2\ C_8H_{19}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(l)$$

From the following table, calculate the total amount of energy available to drive a 1500 kg car by burning 40.0 L of gasoline, assuming 25% efficiency of converting the chemical energy in the fuel into mechanical (kinetic) energy of the car. The density of n-octane is 703 kg/m^3.

Compound	**ΔH_f (kJ/mol)**	**MW (g/mol)**
C_8H_{18}(l)	-208.27	114.2285
O_2(g)	0	31.9988
CO_2(g)	-393.475	44.0095
H_2O(l)	-285.828	18.01528

First, let's calculate the enthalpy of combustion for octane from the data provided.

$$\Delta H_c = (16\ mol)\left(-393.457\frac{kJ}{mol}\right) + (18\ mol)\left(-285.828\frac{kJ}{mol}\right) - (2\ mol)\left(-208.27\frac{kJ}{mol}\right)$$
$$= -11023.6\underline{7}6\ kJ$$

This means that for every 2 mol of C_8H_{18}(l) combusted, 11023.6$\underline{7}$6 kJ of energy will be released. Now, we just need to figure out how many mol of C_8H_{18}(l) there are in 40.0 L.

$$40.0\ L \cdot \frac{m^3}{1000\ L} \cdot \frac{703\ kg}{m^3} \cdot \frac{1000\ g}{kg} \cdot \frac{mol\ C_8H_{18}}{114.2285\ g} = 24\underline{6}.173\ mol\ C_8H_{18}$$

So, the total energy available, assuming 25% efficiency is given by

$$\left(24\underline{6}.173\ mol \cdot \frac{11023.6\underline{7}6\ kJ}{2\ mol}\right)\left(\frac{25\ \%}{100\ \%}\right) = 33\underline{9},216.424\ kJ$$